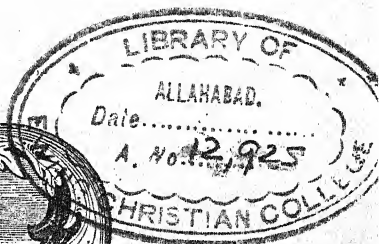
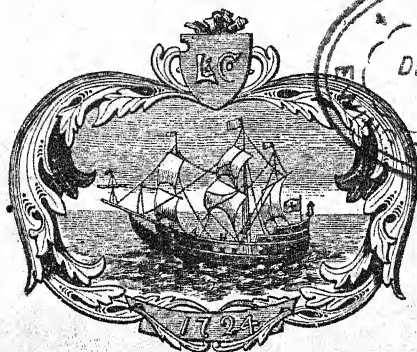


MODERN INORGANIC CHEMISTRY

BY

J. W. MELLOR, D.Sc., F.R.S.



WITH ILLUSTRATIONS

NEW IMPRESSION

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PREFACE

The true aim of the teacher must be to impart an appreciation of method, and not a knowledge of facts.—K. PEARSON.

The power to recognize and to follow truth cannot be conferred by academical degrees.—PARACELSUS.

EVERY teacher now recognizes that it is a sheer waste of time to introduce many abstract ideas into an elementary science course without a previous survey of facts from which the generalizations can be derived. In most cases the historical mode of treatment is correct, because the generalizations have usually developed from a contemplation of the facts; in other cases the historical treatment may involve digressions which would seriously interfere with the efficiency of the course. Obviously, a teacher will try his best to instil the maximum amount of scientific method into the facts—as prescribed by his syllabus and time-table—always remembering that the student gets more lasting benefit from the method than from the facts *per se*. In after-life the scientific method may be retained as a permanent attitude of the mind when the facts themselves are nearly all forgotten. Consequently, the teacher seeks to develop a certain spirit or attitude of mind which is almost equivalent to a sixth sense, and therefore the justification for a general course in chemistry must be sought in the mind of the student rather than in the facts of the science. What, then, may a student expect from a general course of chemistry?

1. *Skill in observation and experiment.*—All are agreed that personal contact with facts is a great advantage. The constant absorption of statements and opinions from text-books makes a student lean so much on authority that he ultimately becomes unfitted for independent observation. Habits of self-reliance, resource, and initiative can be acquired only in the laboratory, or by direct contact with the facts themselves. But practice in observation and experiment is not alone sufficient to develop the scientific faculty. The observational powers of a savage are usually keener than those of a civilized man, and a student may learn to observe without gaining much beyond an increased facility in the art: and he may

become very skilful in experimenting without gaining much more than mere dexterity in manipulation.

2. *Memory and knowledge of relevant facts.*—Facts, of course, form the raw material which is refined by scientific methods into science itself. Science can do nothing without facts. Consequently, many facts must be memorized by the neophyte in chemistry. Some students soon learn the trick of amassing and memorizing all kinds of information in a mechanical way. All the facts associated with a phenomenon may not be of equal importance. In practice it is not always easy to discriminate between relevant and irrelevant facts. Still, it is important to confine the attention as closely as possible to relevant and essential facts, and to discard those irrelevant and accidental. The tyro in chemistry must trust his teacher to indicate the more significant facts to be committed to memory, and used as material for exercising his intellect and wits.

3. *Ability to reason and think in a logical systematic way.*—A student must learn to reflect on the available data bearing on the problem in hand, and to explain a phenomenon by drawing legitimate inferences from approved evidence. It is a mistake to postpone the exercise and discipline of the thinking faculties until a student has memorized a vast accumulation of facts. It is necessary to form habits of reflection and thought as early as possible. Exercise means growth. It is far easier to acquire a mass of facts than to learn to draw a sound inference as to what the facts prove. The thinking faculty can be developed only through the student's own individual efforts. Just as the memory, in some subtle way, grows more vigorous with use, so the exercise of the thinking faculties enhances the power to think. Every exercise of the reason, said Sir Humphry Davy, in 1811, strengthens the habit of correct thinking, and adds sometimes to the influence and power of common sense. Vague indefinite observing is usually followed by muddled inchoate thinking. Clear thinking pre-supposes clear seeing.

4. *Cultivation of the imagination.*—Some teachers have very pronounced objections to the introduction of scientific theories in an elementary course; they claim that "it is not scientific to present and discuss, say, the atomic theory in an elementary chemistry course." It might be asked what constitutes an elementary course? It would be a great mistake to suppose that science has no need for the imagination, for it is very true, as K. Pearson has said, that "disciplined imagination has been at the bottom of all great scientific discoveries"; and, as W. A. Fiske has said, that "every hypothesis and law of science is the result of a vivid imagination." Imagination helps to complete the picture outlined by observation and inference. The picture must, of course, be tested and criticized in every con-

ceivable way to make sure that it is not a mirage among the purpled morning clouds to be dispelled by the dawning light.

5. *Development of a critical and impartial judgment.*—The imagination, though very useful, is a most dangerous ally; and a sharp line of demarcation must be observed between valid or legitimate deductions from the evidence and what has been supplied by the imagination. Each proposition must be judged solely on its merits. There must be no shirking of the facts, no exaggeration, no distortion of the naked truth. The mind must be kept open and free from prejudice. The student must learn not to prejudge data and phenomena by ideas formed independently of the things themselves. A teacher soon accumulates remarkable examples of the influence of expectation on judgment. If a practical class knows what quantitative result "ought" to be obtained, it is surprising how much nearer that result the majority will get than if the true result were unknown—and this without dishonest intention. Rigorous honesty and absolute impartiality in dealing with approved evidence are indispensable. A complete absence of bias can alone give reality and meaning to scientific truths.

I have to thank several authorities for permission to use a number of quotations and a selection of questions from college examination papers. The source of each is indicated in the text. The original wording of the examination questions has been slightly modified in a very few cases. I have pleasure in thanking a number of friends for reading portions of the proofs. In some cases I have persuaded friends who have specially studied a particular phenomenon or process to glance through the proofs to make sure all is sound. I gratefully acknowledge the help I have consciously and sub-consciously received from the examination and review of a large number of text-books during the past few years. I have been very fortunate in having had the difficulties experienced by students with the old edition placed before me by a number of teachers, and I have taken advantage of their kindness to incorporate many of the results in the new text. For this they have my best thanks. I have been asked at various times for the authority of many statements; these, accompanied usually by fuller details, will be found in *A Comprehensive Treatise on Inorganic and Theoretical Chemistry*. I am greatly indebted to Messrs. H. Edwards, G. White, H. Oliver, B.Sc., and A. Johnson, for reading the proofs.

J. W. M.

SANDON HOUSE,
STOKE-ON-TRENT.

February, 1925. •



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MODERN INORGANIC CHEMISTRY

CHAPTER I

INTRODUCTION

§ 1. The Evolution of Chemistry.

The progress of science is as orderly and determinate as the movement of the planets, the solar systems, and the celestial firmaments. It is regulated by laws as exact and irresistible as those of astronomy, optics, and chemistry.—S. BROWN (1843).

As a result of fiery controversies waged about the philosophy of Aristotle, we are told that the citizens of Geneva, towards the end of the 16th century, decreed: "For once and for ever, in no branch of learning shall any one stray from the philosophy of Aristotle." The modern student cannot obey this dictum in blind faith. Knowledge grows. There is a universal law of progressive development whereby error gradually diminishes and truth expands. Consequently, the opinions of Aristotle were founded upon a less developed and a cruder knowledge of facts and phenomena than is the case with the present generation, which has a wonderful heritage of accumulated experience and ideas which have been stored, selected, and sifted by centuries of toil and labour.

T. Bergman, and A. Comte have shown that it is characteristic of the progressing human mind to employ three different methods of explaining natural phenomena, and that man's opinions concerning natural processes can usually be traced through three epochs: (i) the mythological or superstitious, (ii) the metaphysical or abstract, and (iii) the positive or scientific. Otherwise expressed, in the first stages of his learning, man attempted to explain natural phenomena by supernatural agencies; in the second, by abstract reasoning; and in the third stage, by ascertaining the laws of succession and similitude by the study of accurately observed facts. This famous hypothesis has been called Comte's law of three states. True, the three epochs may overlap and blend with one another so that no sharp boundaries can be made with respect to time or place, and, owing to special circumstances, there may be a retrograde movement by the advent of an age of intellectual darkness; yet, in the main, these three periods characterize the growth of science as surely as

the child, the boy, and the man characterize the development of an individual's mind. Chemistry is a particularly happy illustration of Comte's idea.

I. The first, the mythological, or superstitious stage.—This represents the childhood of chemistry, for, as man emerged from the mists of prehistoric antiquity, everything must have appeared to be full of wonder and mystery. He was overawed by the wind and the rain; by the lightning and the thunder; by the eclipse and the comet; and by the rainbow and the clouds. The student of nature lived in a bewildering dreamland of mixed magic and myth which led him to ascribe supernatural explanations to inaccurately known facts, and consequently, he seemed to be surrounded on all sides by an invisible world of spirits and demons. Just as man's own actions seemed to be the result of his own efforts and volitions, so did natural phenomena appear to be the work of benignant or malignant spirits in air, earth, or sea; and man accordingly made oblations to their residing deities to secure their kindly offices. Chemical phenomena were produced by spirits—the salamander or the sylph, the naiad or the nymph, the undine or the gnome—indwelling in different bodies, whose aid was invoked by incantation or charm to produce successful experiments.

Accordingly, men who studied nature in those days were often suspected of tampering with the spirits of evil, and chemistry came to be known as one of the seven devilish arts. So too arose a childish fear and hatred of science, and the belief—widespread in the Middle Ages—that science is dangerous, and its votaries ought to be suppressed. In illustration, in 1287 it was proposed to suppress chemical studies as had been attempted with physics in 1243; again, the *Accademia dei Segreti*—Academy of Nature's Secrets—founded by J. B. Porta in 1602 for the discussion of scientific subjects, was dissolved after a short but glorious existence, apparently because it was believed that magic and the black arts were practised at its meetings. In the 13th century, Roger Bacon was arraigned at Oxford on an indictment for practising sorcery and magic; and in order to disprove these accusations, he wrote his celebrated *Epistola*, which showed that phenomena and appearances, then attributed to supernatural agencies, were simply due to the operation of natural laws. Again, in the middle of the 16th century, J. B. Porta tried to show that the magic of nature is quite as wonderful as that of wizards and witches.

The cuneiform inscriptions and the records of antiquity which have been transmitted to us, show that the early chemists were dominated by the gratuitous assumption that "the interior agencies which keep the world in motion were personal forces essentially out of and above nature." The magician and the sorcerer, the necromancer and the wizard, were the founders and keepers of the first rudimentary knowledge of nature. Accordingly, knowledge and superstition were interwoven with wondrous ingenuity and subtlety. The alchemists, following the mysticism introduced by the Alexandrian and Arabian schools, had virtually reverted to this stage of development when they spoke of red bridegrooms (gold) and lily brides (silver); of green dragons (mercury) and red lions (gold); of black crows (lead) and yellow scorpions (sulphur); and of flying eagles, fugitive stags, and inflated toads. One of the older chemists described the result of triturating mercuric chloride with mercury, resulting in the

formation of mercurous chloride, in these pompous words: "The fierce serpent is tamed, and the dragon so reduced to subjection as to oblige him to devour his own tail." Another old book represented the dissolution of gold in aqua regia by a lion devouring the sun as depicted in Fig. 1. The seven metals—gold, silver, an alloy, copper, tin, iron, and lead—known to the early Chaldeans, were also designated by the names and symbols of the seven greater heavenly bodies—the Sun, Moon, Mercury, Venus, Jupiter, Mars, and Saturn. A close relation was supposed to subsist between the metals and their respective planets, so that nothing could happen to the one which was not shared by the other; and it was further supposed that experiments with any particular metal were more likely to succeed when the governing planet was in the ascendant, and near its zenith.



FIG. 1.—An Old Symbol representing the Dissolution of Gold in Aqua Regia.

II. The second or philosophical stage.—At last man roused himself from his stupor of helpless wonder and childish guessing. He dimly realized some method in nature's inscrutable complexity. Unfortunately, his vision was soon bedimmed and his mind intoxicated. Accordingly, we now find him arrogantly proclaiming the supremacy and omnipotence of the human reason. The majority of educated people of that age believed it to be undignified for a self-respecting man to make experiments, and they did not consider knowledge obtained by observing nature to be a serious subject worthy of mental occupation. Indeed, men were so proud of their intellectual supremacy that they persuaded themselves that their fancies about nature were finer, nobler, and more worthy of belief than nature herself; and Plato apparently considered that the secret laws of nature could be invented by abstract thinking; for he said that "real knowledge is obtained by a simple process of reasoning independently of all information furnished by the senses"; and he expressed his delight with the saying that "the mind is the cause and orderer of all things." The numerous absurdities obtained by the application of this principle emphasize the vanity of the attempt to explain incomprehensible facts by nebulous words; for example, Plato stated:

The universe is a unique, perfect, and spherical production, because the sphere is the most perfect of figures; and it is animated and endowed with reason, because that which is animated and endowed with reason is better than that which is not.

Even Aristotle, the father of logic, reasoned that a vessel containing ashes would hold as much water as when the vessel contained no ashes. The

conclusion is not true, showing that Aristotle did not always recognize the need for the discipline of the imagination by relentlessly checking reason against inexorable fact.

Thus, man did not always see with Cicero that nature is a better teacher than the most ingenious philosopher. The methods of thinking, the much-vaunted philosophy of Plato and Socrates, in its attempt to proclaim the laws of nature from the throne of human reason, actually obscured the path of progress for many centuries, for it became the fashion to look with lofty scorn on knowledge gleaned by observing nature. Accordingly, the leading philosophers worshipped what Francis Bacon might have called idols of the imagination; they devoted themselves to fantastic and chimerical hypotheses about material things; and made no earnest attempt to discriminate between the unreal and the real. As a result, their minds became so prejudiced that the facts were either denied, or else explained by extravagant ideas and fancies uncontrolled by truth and reality as we understand these terms to-day.

III. The third, the scientific, or the positive era.—The marvellous Greeks gave promise of inaugurating this era before the advent of Christianity, but the feeble light kindled by Aristotle flickered and almost expired in the atmosphere of mysticism which prevailed in the Middle Ages. During this period, man almost reverted to the pandemonium of miracle and magic of his childhood days. The light reappeared about the 13th century, and gained brilliancy during the succeeding centuries; man then learned to see that nature is as she is, and is not subjected to the capricious will of deity or demon; man recognized that nature is always conformable with herself without contradictions and without inconsistencies.

The evolution of this stage of chemistry, physics, etc., is connected with the development of the so-called *scientific societies*. The *Accademia del Cimento*, founded at Florence in 1657, under the presidency of Prince Leopold de Medici, was the first scientific society of any importance; its main object was "the repudiation of any favourite system or sect of philosophy, and the obligation to investigate nature by the pure light of experiment." Although it lived but ten years, it enriched the world by leaving a volume of important records of experiments, chiefly in pneumatics. The *Royal Society* of London was founded in 1660; the *Académie des Sciences* of Paris in 1666; the *Academia naturæ curiosorum* of Germany in 1752; and many others were founded in the 18th century. In some exceptional cases, these associations degenerated into "fastnesses from which prejudice and error were latest in being expelled; and they joined in persecuting the reformers of science." In general, however, the policy of these associations was to encourage the investigation of nature by observation and experiment; *Arrière les théories, vivent les faits!* was their watchword; and, instead of clothing their results in enigmatical and allegorical language, they sought to give a candid and straightforward account of their investigations and thoughts. In this way, the obscure mysticism of the Middle Ages was gradually dispelled. Man rediscovered that he does not bring any knowledge into the world with him; that "the subtleties of nature far transcend the subtleties of the human reason"; and that "knowledge cannot be invented, it must be discovered." Progress was then assured, and the manifold achievements of the observational

and positive sciences during the past century are in striking contrast with the paucity of the results of philosophical thinking applied in vain for thousands of years.

§ 2. The Observation and Record of Facts. Collecting Data.

The mind is like a blank tablet upon which experience writes that which is perceived by the senses.—ARISTOTLE (320 B.C.).

To what can we refer for knowledge? What can be a more certain criterion than the senses themselves? If we cannot trust the senses, how is it possible to distinguish what is true from what is false?—LUCRETIVS (60 B.C.).

I know only that truth is in the things and not in my mind which judges them, and that the less I put my mind in my judgments about them, the more sure am I to come near to the truth.—J. J. ROUSSEAU (1770).

H. Poincaré has emphasized in a very telling manner that true knowledge about material things can be acquired only through the senses—*experientia docet*; there is no other way. Experience is the well-spring of true knowledge; experience alone can teach something new; it alone is irrefutable; it alone can give certainty. The same idea was suggested by Aristotle and the peripatetical philosophers: "nothing is in the intellect which was not first in the senses," and by Roger Bacon about 1266. Experience comprises all the impressions we observe and perceive through the various organs of sense. These impressions are recorded in our notebooks, dictionaries of chemistry, etc., as empirical realities or facts. Although knowledge cannot transcend the human faculties, much of the data of science is not *directly* furnished by the senses, for the senses are quite unable to discriminate the subtleties of nature. For instance, the speed of light and the size of atoms are magnitudes either too great or too small to be accessible to sense perceptions. Yet much data derived indirectly from the insensible physical world are assumed to be realities or facts, when actually they are known only by inference from data furnished by the senses. Without facts, science can do nothing; they are the foundation and building stones of the whole superstructure. The edifice can be stable only in so far as it is founded upon the immutability of facts. The facts must be accurate, or the edifice will be unstable.

Not very many years ago, an apt quotation from one of the classical writers—say Aristotle—was considered ample proof of the truth of any statement. Science does not accept P. Bonus' dictum: "The mere fact that a great body of learned men believe a statement supersedes the necessity for proof." To-day, science looks askance on records of mere opinions, and focuses its attention on records of facts. It is not always easy to record facts faithfully without unconscious distortion or bias. What we wish, said Demosthenes, that we believe; what we expect, said Aristotle, that we find. Things are not always what they seem. Seeing is not always believing. It is often difficult to distinguish appearances from realities, for we are easily deceived by the mockery of sensations. The senses cannot be divorced from the mind; neither is always to be trusted alone. The sun appears to rise and set; in reality it does neither. So, although experience is the source of truth, it may also be a source of error. Superficial appearances may obscure hidden realities. Plato of

old was unduly oppressed with the illusions and deceptions of sensory impressions, and he was accordingly led to deny the validity of knowledge derived from the sensations; but Aristotle rightly showed that difficulties arise only when the mind wrongly interprets the testimony of the senses.

It is therefore sometimes necessary to receive with caution the testimony of evidence derived from sensations. The mind interprets a sensation by comparing it with some former sensation, the source of which has been previously determined. Consequently, the faithfulness of the interpretation is dependent upon the memory of past sensations, or upon the sensitiveness of the mind to detect resemblances and differences. Otherwise expressed, the accuracy of an inference as to the nature of the *objective* source of a *subjective* sensation varies from a mere guess to virtual certainty. The idea has been aptly illustrated this wise: just as a number of bits of glass irregularly arranged always form symmetrical patterns when viewed through the kaleidoscope, so does the understanding of each man impose a pattern of its own upon the various sensations which it perceives. Consequently, as Robert Hooke once said: It is necessary to be on guard against deep-rooted errors which may have been grafted upon science by the slipperiness of the memory, the narrowness of the senses, and the rashness of the understanding. The greatest caution must be exercised in accepting, on secondhand evidence, facts which cannot be verified. No reliance can be placed on vague impressions. Evidence must be clear and precise.

Few persons can estimate and register facts impartially and fairly. As W. S. Jevons puts it: "Among uncultured observers, the tendency to remark favourable and forget unfavourable events is so great that no reliance can be placed on their supposed observations." Untutored minds are very prone to mistake inferences for observations, and prepossessions for facts; their observations and their judgments are alike vitiated by dogma and prejudice; they do not seek to investigate, they seek to prove. The old proverb is inverted, believing is seeing. The student of science must pledge himself to do his best to eliminate prepossession and dogma from his judgments, and he must spare no pains to acquire the habit of recording phenomena as they are observed; and to distinguish sharply between what is or has been actually seen, and what is mentally supplied. It requires a mind disciplined like a soldier to avoid the natural inclination to look away from unwelcome facts.

The purity of truth is almost certain to be corrupted when the observer is dominated by preconceived opinions, for, as O. W. Holmes puts it: When we have found one fact, we are very apt to supply the next out of the imagination. It is only in a pseudo-science that positive evidence, or such as tells in favour of its doctrines, is admitted; and all negative evidence, or such as tells against it, is excluded. C. Darwin has stated that one of his golden rules was to make a memorandum of any fact or thought which he found to oppose his general results, because he noticed by experience that such facts or thoughts were far more apt to escape the memory than favourable ones. Above all, said Robert Hooke, a good observer needs a sincere hand and a faithful eye, to examine and record things themselves as they really appear. "The mind and the reason of the trustworthy observer must be trained to rebel against all desire, and to disobey all inclinations."

§ 3. The Collating, Sifting, and Clarifying of Observations. Classifying Data.

History teaches that the commencement of every branch of science is nothing more than a series of observations and experiments which had no obvious connection with one another.—J. VON LIEBIG (1846).

In order that the facts obtained by observation and experiment may be capable of being used in furtherance of our exact and solid knowledge, they must be apprehended and analysed according to some conceptions which, applied for this purpose, give distinct and definite results, such as can be steadily taken hold of, and reasoned from.—W. WHEWELL.

The record of facts obtained by observation and experiment, *per se*, is empirical knowledge. *Empirical* is derived from a Greek word meaning experienced. It has just been emphasized that all knowledge is derived from experience, and hence empiricism would appear to be the right method of acquiring knowledge. The term, however, has slightly changed in meaning, for it is now usually applied to chance experiences which occur irregularly without any orderly plan of investigation.

All true science, said T. Huxley, must begin with empirical knowledge. Nature, however, presents to our senses a panorama of phenomena co-mingled in endless variety so that we are sometimes overwhelmed and dazed by the apparent complexity of empirical knowledge. It is work for the intellect to educe the elements of sameness amidst apparent diversity, and to see differences amidst apparent identity. It is work for the judgment to reject accidental and transient attributes, and to consolidate essential and abiding qualities. Consequently, while the primary aim of science is to collect facts, the higher purpose of science is to show that, amidst wild and terrible disorder, order and law reign supreme. The man of science seeks a refuge from this bewildering complexity in unifying principles by which the facts can be grouped and classified into systems. As he gazes into nature, the man of science must be quick to discern hidden resemblances amidst a thousand differences; he must be quick to disentangle natural relations from a medley of detail; and quick to detect dissemblances amidst alluring similarities.

Empirical knowledge describes facts; science begins by comparing facts. Empirical facts, in consequence, can form a science only when they have been arranged, rearranged, grouped, or classified so as to emphasize the elements of similarity and identity in different phenomena. Accordingly Thomas Hobbes expressed the opinion that the main purpose of science is the tying of facts into bundles. This bundle-tying, indeed, forms no small or insignificant part in the development of science; otherwise expressed, a significant advance has been made in the development of a science when the observed facts have been codified into a system so that a medley of empirical facts is systematically summarized under a small number of heads. This means that the facts must be arranged in a methodical and systematic manner until finally all the relevant facts taken together may form one system. The process of classification and correlation is one of the methods of scientific investigation. Knowledge so systematized is scientific knowledge. T. Bergman illustrated the idea by saying that "a vast number of observations without order or regularity

is not unlike a confused heap of stones, lime, beams, and rafters requisite for constructing an edifice, but which being combined with no skill fail in producing the proposed effect."

The material framework of the world appears in a myriad different guises and combinations, but the chemist can resolve each combination into a few definite elementary forms of matter; similarly, a multitude of forces can be resolved into comparatively a few primitive forms of energy. About 150 A.D., the Egyptian astronomer Claudius Ptolemy measured the angles of incidence and refraction of a beam of light passing from air into water, but more than fourteen hundred years elapsed before W. Snell detected the law of refraction hidden in Ptolemy's data. By tabulating his measurements of the volumes of air confined under different pressures, Robert Boyle discovered the law known by his name. Each of these laws summarizes in one simple rule myriads of possible measurements.

Scientific knowledge is not necessarily more accurate than empirical knowledge. Empirical unco-ordinated facts are no less true, definite, and real than scientific facts, for all facts are equally true *per se*. A collection of empirical facts always requires some theory to serve as framework in order that the facts may be arranged, grouped, and pigeon-holed. If a group of facts—scientific facts—has been organized on an erroneous system, the facts are no less true though the system be false. Chemistry presents a curious mixture of empirical facts with isolated fragments of scientific knowledge.

§ 4. The Generalization of Observations.

Facts are the body of science, and the idea of those facts is its spirit.—
S. BROWN.

It is the intuition of unity amid diversity which impels the mind to form science.—F. S. HOFFMAN.

The correlation of empirical facts requires qualities of the mind different from those employed in observation and experiment. Both qualities are not always located in the same individual. Some excel in the one, not in the other. J. Priestley, C. W. Scheele, and H. Davy, for instance, were admirable observers, but they were not brilliant in the work of correlation; J. Dalton, and A. L. Lavoisier were not particularly distinguished as experimenters, but they excelled in correlating observed data. W. Hamilton did not rate the fact-collecting faculty very highly. He said that "in physical science the discovery of new facts is open to every block-head with patience, manual dexterity, and acute senses; it is less effectively promoted by genius than by co-operation, and more frequently the result of accident than of design." J. Priestley recognized his own limitations when he said: "I have a tolerably good habit of circumspection with respect to facts, but as to conclusions from them, I am not apt to be very confident." Skill in the critical analysis of observational data, and in collating, sifting, and clarifying records, is not a sufficient recommendation to the adytum—the *sanctorum sanctissimum*—of science. There is still a higher type of work for but a few seekers after knowledge. It is

To search thro' all
And reach the law within the law.—TENNYSON.

It is the sprite imagination which usually reveals the deeper meaning of facts which have been diligently garnered and laboriously sifted. It cannot be doubted that science in its higher work requires a supple and well-developed imagination which T. Gomperz says is the instrument of genius, no less for scientific discovery than for artistic creation. The secret charm of scientific discovery is not in the facts *per se*, but rather in the extrication of natural relations among the facts one with another. Particular groups of facts must be unified or generalized into a system—the so-called law. Science begins with facts and ends with laws. Law is the essence of facts. As pointed out elsewhere, Newton's celebrated law epitomizes in one simple statement how bodies have always been observed to fall in the past. Immortal Newton did not discover the cause or the why of the falling of the apple, but he did show that it was due to the operation of the same forces which hold the earth, the planets, and their satellites in their appropriate orbits. Newton's simple and comprehensive law epitomizes in one single principle the many and varied phenomena associated with falling bodies, planetary motions, etc., and generally, the scientific generalization explains the operations of nature by showing the elements of sameness in what at first sight appears to be a confused jumble of phenomena. Generalization is the golden thread which binds many facts into one simple description. That peculiar type of genius, that rare quality of mind required for the work of generalization, is found only in a Newton or a Darwin. Plato said that if ever he found a man who could detect the *one in many* he would follow him as a god. Unification is the supreme goal of modern science, or, as Heraclitus (c. 450 B.C.) proclaimed, the highest goal of knowledge is the one law regulating all events. At best, man has to apply a very weak intellect to a very complicated world; and the resources of the human intellect are too narrow, and the universe is too complex to leave any hope that it will ever be within man's power to carry scientific perfection to Tennyson's last degree of simplicity:

. . . one law, one element.

§ 5. The Aim of Science in General, and of Chemistry in Particular.

Let us remember, please, that the search for the constitution of the world is one of the greatest and noblest problems presented by nature.—G. GALILEI.

The ordered beauty of the world of nature suggests an infinite intelligence with powers of action such as no man possesses.—BENJAMIN MOORE.

Science embraces the sum-total of human knowledge, and it ranges over the whole realm of nature. Science is not a mass of empirical knowledge gained by observation and experiment, but it is an organized body of facts which have been co-ordinated and generalized into a system. Science tacitly assumes that nature is a harmonious unity, and that rational order pervades the universe. Science seeks a complete knowledge of the multitude of inter-related parts of the universe which act and react on one another producing endless variety. In fine, science aims at omniscience. The target, however, appears to recede with increasing knowledge. As man grows in wisdom and knowledge, he begins dimly

to realize that the unknown multiplies into boundless proportions, and science might well confess with Tennyson :

So runs my dream. But what am I ?
An infant crying in the night ;
An infant crying for the light ;
And with no language but a cry.

The sciences are too complex and too vast to be comprehended by one man's mind.

One science only will one genius fit,
So vast is art, so narrow human wit.—POPE.

Our feeble wit has rendered it necessary to rear a tree of scientific knowledge with many branches : astronomy, physics, chemistry, mineralogy, geology, biology, sociology, etc. "The divisions of the sciences," said Francis Bacon, "are like the branches of a tree that join in one trunk," and therefore they are more or less closely related with one another. The astronomer, the physicist, the chemist, each usually keeps to his own particular branch. This separation of the sciences is mere convention. Even in the middle of the 13th century Roger Bacon saw that there are no real lines of demarcation between the different sciences, for he pointed out that "all the sciences are connected ; they lend each other material aid as parts of one great whole. Each does its own work, not for itself alone, but for the other parts. . . . No part can attain its proper result separately ; since all are parts of one and the same complete wisdom."

The science of chemistry is man's attempt to classify his knowledge of all the different kinds of matter in the universe ; of the ultimate constitution of matter ; and of the phenomena which occur when the different kinds of matter react one with another. The science of chemistry is itself so vast, that many branchlets are necessary for useful work, and thus we have : inorganic chemistry, organic chemistry, physical chemistry, mineralogical chemistry, bio-chemistry, agricultural chemistry, pharmaceutical chemistry, etc. The chemist also frequently aims at applying his knowledge to useful purposes in the arts and industries ; and thus arises *applied, industrial, or technical chemistry*.

About the middle of the 13th century, Roger Bacon distinguished between knowledge sought for the sake of truth, and knowledge utilized in the practice of the various arts ; or, as I. R. Averroës expressed it a century earlier : In pure science, we learn so that we may know ; and in applied science, we learn so that we may apply our knowledge to useful purposes. The distinction, however, was recognized in the 4th century B.C., for it was explicitly expounded by Aristotle, and by Plato. The purpose of pure science is to observe phenomena and to trace their laws ; the purpose of art is to produce, modify, or destroy. Strictly speaking there is no such thing as applied science, for, the moment the attempt is made to apply, science passes into the realm of art. It has been well said that "science is indebted to art for the means of experimenting, but she instructs art concerning the properties and laws of the materials upon which the latter operates." In an essay on *The usefulness of experimental philosophy*, Robert Boyle emphasized the mutual benefits which would obtain when science, or, as he called it, when natural philosophy is applied

to the various arts and crafts ; and he claimed that it is prejudice, no less pernicious than general, which has kept science so long a stranger in the industries. Boyle's ideas have been still further emphasized by Lord Kelvin, who said :

There cannot be a greater mistake than looking superciliously upon practical applications of science. The life and soul of science is its practical application, and just as the great advances in mathematics have been made through the desire of discovering the solutions of problems which were of a highly practical kind in mathematical science, so in physical science many of the greatest advances that have been made from the beginning of the world to the present time have been in the earnest desire to turn the knowledge of the properties of matter to some purpose useful to mankind.

The so-called applications of science to the industrial arts—say, applied chemistry—may be (i) An attempt to extend the methods of scientific investigation to the industrial arts ; or (ii) To adapt known operations and laws to useful purposes. When the chemist is occupied in the systematic observation of phenomena, and in tracing their laws, he is engaged in scientific investigation, no matter if the work be conducted in academy, in counting house, or in factory.

§ 6. Experiment.

Experiment is the interpreter of nature. Experiments never deceive. It is our judgment which sometimes deceives itself because it expects results which experiment refuses. We must consult experiment, varying the circumstances, until we have deduced general rules, for experiment alone can furnish reliable rules.—LEONARDO DA VINCI.

Nature speaks to us in a peculiar language, the language of phenomena. She answers all the questions we ask her, and these questions are our experiments.—J. VON LIEBIG.

Chemistry is largely an experimental science. Experiment is really a method of observation, which is employed when the facts are so masked by other conditions that they cannot be accurately observed unless the obscuring conditions are suppressed. The chemist would not make much progress if it were only possible to observe phenomena just as they occur in nature, and not possible to make observations under determinate conditions. By experiment, it is possible to make combinations of different forces, and different forms of matter which are not known to occur in nature ; to eliminate complex disturbing conditions ; and to observe phenomena under simplified conditions. An experiment has been well defined as *une observation provoquée*. Experiment is useful only when there are conditions which obscure direct observations. The most successful experiment does no more than make a fact which was previously obscure as patent as one that was open to direct observation from the first. Chemical phenomena, *per se*, are usually too complex for our minds to grapple, and they must be simplified by simple experiments. Consequently, chemistry is an experimental science because its facts can rarely be observed in any other way. If data could be obtained by direct observation, there would be no need for experiment.

It requires much acumen to determine the precise conditions under which an experiment shall give a successful result. Every experiment has the character of a specific question. The skilled questioner—the experimenter—knows what he is asking, and he tries his best to interpret

nature's reply, be it affirmative, negative, or evasive. If the answer be negative or evasive, the question has not been properly asked, and it must be plied again and again until

A sharphooked question baited with such skill
It needs must catch the answer.

Paradoxically enough, the investigator can usually say with "Dr. Moreau": "I asked a question, devised some method, of getting an answer, and got—a fresh question." Some such ideas were in Robert Hooke's mind when he said: "The footsteps of nature are to be traced, not only in her ordinary course, but when she seems to be put to her shifts, to make doublings, and turnings, and to use some kind of art in endeavouring to avoid our discovery."

The more intricate the experiment, the greater the probability of an obscure and ambiguous result. As A. L. Lavoisier has pointed out, "it is a necessary principle in experimental work to eliminate every complication, and to make experiments as simple as possible." The *quality* of an experiment, not the *quantity*, is best adapted to throw light upon a phenomenon. Experiments carelessly performed may be sources of error and obscurity. Many of the results obtained by the alchemists in the Middle Ages show how ineffective or abortive are the results of experiments in incompetent hands—here, the experiments wandered into eccentric by-paths, and furnished preposterous conclusions. Experiment is an art, said G. A. Lewes, and demands an artist.

Joseph Priestley believed in making a large number of haphazard experiments, and said that he discovered oxygen by trying the effect of heat on many substances, apparently selected at random by John Warltire of Birmingham. Thomas A. Edison, also, appears to have discovered the phosphorescence of calcium tungstate when exposed to Röntgen's rays by deliberately trying the effects of these rays on a large collection of different substances. This old prosaic method of experimenting by trying everything is necessary in some cases, and, though usually dubbed empirical or rule-of-thumb, the process is fundamentally scientific, but it is not generally economical in time and labour. Discoveries are then due, as J. Priestley once argued, more to "chance than to any proper design or preconceived theory." More frequently, the track of the experimenter is blazed by means of working hypotheses.

§ 7. Hypothesis, Theory, and Law.

We are gifted with the power of imagination, and by this power we can enlighten the darkness which surrounds the world of senses. Bounded and conditioned by co-operant reason, imagination becomes the mightiest instrument of the physical discoverer.—J. TYNDALL.

The nearer to the practical men keep, the mightier their power. The theorist who dreams a rainbow dream, and calls his hypothesis true science, at best is but a paper financier who palms his specious promises for gold.—T. L. HARRIS.

Hypotheses are cradle songs by which the teacher lulls his pupils to sleep.—J. W. GOETHE.

It is a popular belief that the aim of science is to explain things; as a matter of fact, the so-called explanations of science do not usually get

much beyond describing the observed facts in the simplest possible terms so as to make their relations with one another clear and intelligible. The description may emphasize the history of a phenomenon, or the conditions under which the phenomenon occurs: In other words, science may explain a phenomenon by describing how one event is determined by an antecedent action—sometimes called a *cause*; and how one particular set of conditions—the cause—can give rise to another set of conditions—the effect. Science explains a phenomenon (the effect) by showing that it is a necessary or rather a probable consequence of another phenomenon (the cause).

Classical scholars tell us that Aristotle has forty-eight, and Plato sixty-four meanings for the word “cause.” The later metaphysicians have also played a game of shuttle-cock with the term. The word “cause” is usually applied to an event, action, or process which “produces” an effect; or, with R. Shute, cause may be regarded as that which the mind selects as a sign of the coming of that other phenomenon which it calls the effect; or conversely, an effect is regarded as something which the mind selects as a sign of the past existence of a cause. There can, therefore, be no cause without an effect, and no effect without a cause. The one presupposes and completes the other. Hence, as P. Carus has observed, the law of causation describes a transformation in which form alone is changed; and consequently, the law of causation is nothing more nor less than another aspect of the famous law of the conservation of matter and energy. The search for the cause of an event is a search for the determining factors which would produce that event. When the cause of an event has been discovered, the event is said to be explained by the cause.

There are certain circumstances or conditions which may exercise, directly or indirectly, a determinative influence on the effect produced by the activity of a cause; and very often certain *conditions* must obtain before an event can occur, thus the temperature of hydrogen must be raised above its ignition point before combustion can ensue. The effect obtained by burning hydrogen is more vigorous if the flame be in oxygen gas than if it be in air. Hence, an atmosphere of oxygen gas is a favourable condition for the combustion of hydrogen; a reduced pressure is a retarding condition because it hinders the speed of combustion and reduces the vigour of the flame. The term “cause” is frequently employed when *reason* is intended. Gravitation is said to be the cause of the falling of a vase from the mantelpiece, whereas the cause of the fall may have really been a push from the elbow. In the former case, the reason why the vase fell downwards is the very same reason why all masses gravitate, and a push was the real cause of the catastrophe. Here the reason of the fall is referred to an inherent quality of bodies, just as the reason why bodies react chemically is explained by investing matter with an inherent quality or *vis occulta*—chemical affinity. If these distinctions be borne in mind, there is no need for confusing cause, reason, and condition, even if one term be used for all three concepts.

The law of continuity—emphasized by G. W. von Leibniz (1687)—assumes that no interruption between cause and event is possible, and that there is a connected chain in the order of natural phenomena so that when several of the links are known, the intermediate links can be inferred. Consequently, men of science assume that each phenomenon is an effect of a previous event, and is itself the cause of a succeeding effect, and that

under like conditions, the same causes produce the same effects. Apart altogether from the question whether or not nature can do precisely the same thing again under precisely similar circumstances as she has done before, the principle of continuity or uniformity assumes that any phenomenon will be repeated if all the preceding phenomena be precisely repeated; otherwise expressed: the same antecedents are invariably accompanied by the same consequents. Hence, it has been said that science does not now seek for the reason or the why of events, but rather for invariable relations between phenomena. The law of causation is taken to *describe* a sequence of changes starting with the cause and ending with the effect. G. Kirchhoff introduced the term *description* as a synonym for *cause* at the very beginning of his *Lectures on Physics*, where he said: "The object of mechanics is to give a complete *description* in the simplest possible manner of such motions as occur in nature."

Although every effect may be traced to a previous event as its cause, in the physical world, phenomena follow one another as links in an unbroken chain of cause and effect. It is soon recognized that the cause of a phenomenon is an effect which itself needs explaining by some ulterior cause, so that causes can be traced backwards in a never-ending chain of events. Owing to the limited range of man's understanding in a world of infinite complexity, we are far, very far, from comprehending the true conditions, the true causes, or the true reasons for natural phenomena.

The mind cannot receive a long series of details without encircling and connecting them by a common bond which is a kind of *mental nexus*; similarly, in the attempt to find the causes of many phenomena, man is compelled to build an imaginary model showing how a given set of conditions—the hypothesis or theory—is always followed by particular effects. A phenomenon is then explained by showing that it is bound to occur by the operation of the set of conditions postulated by the hypothesis. Consequently, hypotheses are essentially guesses at truth. The rational observer does not trust to random guesses, but he is guided by a more or less vague intuitive conjecture (hypothesis) as to the meaning of the phenomena under investigation, and experiments are devised accordingly, for

Man's work must ever end in failure,
Unless it bear the stamp of mind.
The head must plan with care and thought,
Before the hand can execute.—SCHILLER.

The Spanish philosopher J. L. Balmés emphasized this same idea when he said:

Although one accepts as a *real* truth the most uncontested and the most certain fact, it remains sterile if *ideal* truths do not fecundate it. . . . To acquire scientific value, the facts, being submitted to reflection, must be impregnated by the mind with the light it lends to necessary truths.

Hypotheses precede observation and prompt experiments, for they indicate the conditions under which the search for new facts is likely to be successful. Hence, when Leonardo da Vinci (c. 1500) said that "hypothesis is the general, and experiments are the soldiers," he probably meant that hypotheses direct or indicate what experiments should be made. Accordingly, hypotheses are indispensable aids in the systematic quest after the secret meaning in nature's deeds. Those who

refuse to go beyond fact, said T. H. Huxley, rarely get as far as fact. It is difficult to believe that so astute an investigator as Joseph Priestley really overlooked this mode of investigation, as might be supposed from some preceding remarks—nor did he. On the contrary, he said :

It is by no means necessary to have just views, and a true hypothesis, *a priori*, in order to make real discoveries. Very lame and imperfect theories are sufficient to suggest useful experiments which serve to connect those theories, and give birth to others more perfect. These then occasion further experiments, which bring us still nearer to the truth, and in this method of approximation, we must be content to proceed, and we ought to think ourselves happy if, in this slow method, we make any real progress.

The many gaps in our knowledge are temporarily bridged by the assumptions called hypotheses. Hypotheses thus help to render intelligible the interrelations between different facts, and they are employed by men of science to extend and deepen their experience by predicting and disclosing new facts ; to correct and purify their knowledge of natural phenomena by eliminating errors and contradictions ; and to systematize their description of facts so as to obtain the greatest control over them with the least possible effort.

The verification of hypotheses.—An hypothesis may seem to be the logical consequence of known facts, or it may be a random flash of the imagination. However probable an hypothesis might appear, both the hypothesis and the logical consequences of the hypothesis must be tested by comparison with facts. Aristotle (c. 320 B.C.) certainly recognized the need for basing reasoning on observed facts, but, as G. H. Lewes has emphasized, Aristotle did not realize the very vital importance of verifying his logic by comparing its conclusions with facts, nor did he recognize that the true purpose of experiment is to verify the accuracy of data and of theoretical conclusions. We are indebted to Roger Bacon (c. 1280), perhaps more than to any other, for first insisting on verification as the essential pre-requisite for every trustworthy conclusion. He said :

Experimental science is the mistress of speculative science. She tests and verifies the conclusions of other sciences. . . . In reasoning we commonly distinguish a sophism from a demonstration by verifying the conclusion through experiment.

Experiments have a way of giving results which differ from those which rigorous logic concluded must occur ; and when the prediction fails, it is necessary to find what has been overlooked. This does not mean that constant verification is needed to establish the validity of the *process* of reasoning, for that may be irreproachable and yet the conclusion may be false because the facts or premises upon which the reasoning was founded may have been interpreted to mean something very different from what actually obtains in nature, or because some unrecognized or undiscovered factor was involved. It is not wise to dogmatize when direct trial is possible : " Do not think," said J. Hunter ; " try."

It has been aptly said that the remarkable discoveries of modern science have been made by invariably sifting the truth through a fine mesh of logical experiment. One of C. Darwin's favourite methods of applying this method was to reason : " If my hypothesis be true, then certain consequences must also be true. Now let us find if they are true " ;

and H. St. C. Deville used to say that there is no need to argue if an experiment can be made. In fine, it is necessary to submit all conjectures to the incorruptible test of fact in order to avoid being seduced by immaterial creations of the imagination. Faith without facts availleth nothing. The *ad experimentum* test must be made with unremitting diligence, rigorously and impartially, without conscious bias. Trial by a combat of wits in disputations has no attraction for the seeker after truth; to him, the appeal to experiment is the last and only test of the merit of an opinion, conjecture, or hypothesis.

If one hypothesis does not fit the facts, it is discarded, and a modification of the old, or totally new hypothesis is tried. Thus, J. Kepler, in his *De motibus stellæ martis* (1608), is said to have made nineteen hypotheses respecting the form of planetary orbits, and to have rejected them one by one until he arrived at that which assumed their orbits to be elliptical. "To try wrong guesses," said W. Whewell, "is apparently the only way to hit the right ones." This method of trial and failure is continued until the golden guess crowns the investigation; but one single real conflict between fact and hypothesis will destroy the most plausible hypothesis. Of fifty hypotheses, only one may prove fruitful; the unsatisfactory ones are weeded out, until that particular one remains which has established its right to live by proving itself useful or by satisfying some need. Quoting M. Faraday:

The world little knows how many of the thoughts and theories which have passed through the mind of a scientific investigator have been crushed in silence and secrecy by his own severe criticism and adverse examination; that in the most successful instances not a tenth of the suggestions, the hopes, the wishes, and the preliminary conclusions have been realized.

This quotation may give a wrong impression, for Michael Faraday displayed consummate skill, not only in framing hypotheses *per se*, but in deducing hypotheses that were worth testing. Without hypotheses, the experimental method may degenerate into empiricism; without experiments, hypotheses may degenerate into speculation.

The promulgation of immature or premature hypotheses without a substantial basis of fact is discouraged by most scientific societies. The celebrated nebular hypothesis was ushered in by P. S. de Laplace (1796) with those misgivings and doubts which must of necessity becripple all hypotheses which are not based upon observation or calculation. An hypothesis may be invaluable when it can be verified or refuted by a definite appeal to observation. If this check be not possible, the imagination riots in the wildest speculations. If the evidence of an alleged phenomenon cannot be tested by verification, it is outside the range of science. A. W. Hofmann is reported to have said that he would readily listen to any suggested hypothesis, but on one condition—that he be also shown a method by which it might be tested. Accordingly, scientific inquiry is limited to such objects and phenomena as admit of direct or indirect observational or experimental verification. On the other hand, science cannot enter into the dark territory beyond the scope of man's faculties, and where verification, direct or indirect, is not possible. A vivid imagination can people this region with phantasms and be deluded with the hallucination that these creatures of the imagination are real, substantial, objective facts. It is now generally recognized that imagination,

uncontrolled by facts, has produced many palsyng superstitions which have blinded and cursed the human race in past times.

Rival hypotheses.—Two or more contradictory hypotheses may be consistent with the facts; both cannot be right. There is then need for an *experimentum crucis*, an experiment which will decide in favour of the one and exclude the other. An hypothesis is supposed to be established when it, and it alone, is in harmony with known facts. The hypothesis then ranks as a theory or law. In the majority of cases, the so-called laws of nature can be regarded as prophecies which, because they have always been fulfilled in the past, are expected to be also fulfilled in innumerable cases in the future. Laws, theories, and hypotheses are all on probation. However successful a theory or law may have been in the past, directly it fails to interpret new discoveries its work is finished, and it must be discarded or modified. However plausible the hypothesis, it must be ever ready for sacrifice on the altar of observation. On account of the unproved assumption embodied in all hypotheses, they are of necessity, transient, fleeting, and less stable than theories; and theories, in turn, are less stable than laws. A theory believed to-day may be abandoned to-morrow. New facts need new laws. An hypothesis is invalid when it fails to unite and co-ordinate facts. All our hypotheses and theories are to be superscribed "subject to revision," for they are continually changing. "Science in making is a battlefield of competing theories," the path of progress is strewn with dying and dead hypotheses. For example, W. Ostwald (1893) claims that the theory of chemical combination is a strange and contradictory conglomerate of the fossil constituents of earlier hypotheses. Science is not a *state*, but is rather a *stage* of progress. Even Isaac Newton's law of gravitation is included in this category; and the astronomer R. S. Ball could say:

When the law of gravitation is spoken of as being universal, we are using language infinitely more general than the facts warrant. At the present moment we know only that gravitation exists to a very small extent in a certain indefinitely small portion of space.

Ever since T. Bergman's time (1779), science has been compared with a building in the course of erection, and scientific hypotheses have been compared with the scaffolds and ladders required by the builder in order to place the stones of experience where they belong. The scaffolding must be rejected when it hinders further developments, and when the purpose for which it was erected has been fulfilled. Accordingly, an hypothesis is not the end, but rather the means of attaining that end. To think otherwise would be to suppose that the builder erects a mansion for the sake of showing off the ladders and scaffolds used in its construction. The imperfect notions and hypotheses of men of science must not be mistaken for descriptions of observed facts. In the *chemica docens* of our schools, the term *science* usually includes both the growing building and the auxiliary scaffolding; otherwise expressed, the term includes the immutable facts, the ephemeral hypotheses, the transient theories, and the more or less incomplete generalizations from observations. The facts alone are certain to endure throughout all time. When S. Brown (1849) inquired: Is it necessary to the nature of a science that it be all true, and that it contains no admixture of error? and answered: By no means!

Otherwise chemistry was no science during the reign of phlogiston, and the Lavoisierian chemistry no science so long as oxygen was taken for the principle of acidity, he included in the term *science* those transient theories which are necessarily employed in the erection of the temple of truth.

Deductive and inductive induction.—The term *induction* is applied by the logician to the quest of science for generalizations, that is, for the *canones* or *universales regulæ* of Roger Bacon. In *deduction*, the attempt is made to widen the bounds of knowledge without stepping outside known facts—the Euclidean method is a good illustration; in *induction*, a leap is taken from the known into the illimitable beyond. Two important methods of induction will be recognized—one may be called the deductive method, the other the inductive method. The former was favoured by Francis Bacon, the latter by Isaac Newton.

1. *Bacon's deductive method*, by what he called the *interpretationes naturæ*. Here the facts are exhaustively classified until the generalization becomes clear. a is M or N , or O , or P , or . . . ; but a is not N , nor O , nor P , nor . . . ; and consequently, a is M . Thus, in the 105th aphorism of his *Novum Organum*, 1620, F. Bacon said :

The induction which is to be available for discovery and demonstration . . . must analyse nature by proper rejections and exclusions; and then, after a sufficient number of negatives, come to a conclusion on the affirmative instances.

The method appears to proceed from known facts to general conclusions, *à particulari ad universale*. It is based on facts already known, and has therefore been called *à priori* reasoning. The method by which Boyle's and Charles' laws were discovered might be cited in illustration of one form of the method of deductive induction.

2. *Newton's inductive method*, by what F. Bacon called the *anticipationes naturæ*. Here the attempt is made to infer the hidden generalization from the consequences of the assumption (hypothesis) what that generalization is. The process is sometimes called *à posteriori* reasoning. This method of investigation was extensively employed with glorious results by Isaac Newton, although it had been advocated by Aristotle over a thousand years earlier. Francis Bacon, indeed, before Newton's time, protested against anticipating nature by hypotheses, but the greatest triumphs of modern science have been won by the application of the Newtonian method, while the Baconian method has been singularly unfruitful. Francis Bacon's failure in the practice of his own method was complete.

The particular form which the Newtonian method takes in science is to devise provisional generalizations called hypotheses or *working hypotheses* to explain facts and phenomena. The appeal is then made to observation and experiment in order to test the validity of the proposed generalization. Examples: The cause of the increase in the weight of metals calcined in air; A. L. Lavoisier's theory of combustion, and his experiments on the transformation of water into earth; J. Mayow's work on combustion; etc. The application of this method of inquiry involves (a) The accumulation of facts by observation and experiment; (b) The employment of the imagination in framing hypotheses to explain the facts; and (c) The appeal to facts to prove or disprove the hypotheses. By this procedure, said W. Whewell, the hypothesis becomes the guide of its former teacher

--observation. There is a kind of cycle from facts to hypothesis, and from hypothesis to facts.

Induction, said Aristotle, does not prove. Newton's phrase: *Hypotheses non fingo*—I do not frame hypotheses—is often quoted to show that he discountenanced the inductive method of scientific investigation. This is based upon a misunderstanding, for I. Newton here referred to hypotheses not suggested by observation. On the contrary, I. Newton's own procedure was to use hypotheses deduced from phenomena similar to the way science uses them to-day. Accordingly I. Newton asserted that "no great discovery was ever made without a bold guess," and his immortal *Philosophiæ naturalis principia mathematica* (London, 1687) is a wonderful record of discoveries made possible only by the exercise of the greatest freedom in the elaboration of hypotheses. Indeed, from the first of his communications to the Royal Society on light to the last revision of his *Principia*, Isaac Newton seems to have been steadily and persistently guessing.

The method of investigation employed in scientific, positive, or modern chemistry thus involves four operations: (i) observation and experiment; (ii) classification and comparison; (iii) deduction, or speculation and hypothesis; (iv) testing and verification. Francis Bacon did not grasp the prime importance of testing his induction by comparison with facts. A. de Morgan (1872) puts this rather cleverly: According to Francis Bacon, facts are used to make theories *from*, and according to Isaac Newton, to try ready-made theories *by*. Chemistry could progress as a science only when this method of investigation was discovered, so that as S. Brown stated in 1843, before discovering chemistry it was necessary to discover the art of discovering chemistry.

§ 8. Rey's Experiments on the Calcination of Metals in Air.

Let all the greatest minds in the world be fused into one mind, and let this great mind strain every nerve beyond its power; let it seek diligently on the earth and in the heavens; let it search every nook and cranny of nature; it will only find the cause of the increased weight of the calcined metal in the air.—JEAN REY (1630).

As early as the 8th century, the Arabian chemist Yeber-Abou-Moussah-Djafer Al-Sofi, commonly called Geber,¹ knew that when metallic lead is calcined in air, the resulting calx is heavier than the original metal.² Towards the end of the 15th century, Eck de Sulzbach recognized the

¹ This may be one of the many discoveries attributed to Geber (died 777) which appear to have wrongfully crept into the Latin translations of his writings in the 12th century; for instance, the discovery of sulphuric acid is generally attributed to Geber, although E. Lippmann (1901) has stated that sulphuric acid was not known to the Arabian writers prior to A.D. 975. There are conflicting views about the translations of Geber's works, and each one has been advocated with the keenness of a political debate where facts are few, and opinions many. One side argues that the 12th-century works are veritable translations of the originals; the other side says that if so, the 12th-century work is related as much to Geber's as, say, the current edition of Gmelin's *Handbuch* is related to the edition of 1817-19. No one's opinion, *per se*, is worth much without evidence that he has the necessary linguistic equipment and opportunities.

² The process of heating a metal in air so as to convert it into a calx, is called *calcination*. The calces are generally equivalent to what the modern chemist calls "metallic oxides."

increase in weight which occurs when silver-amalgam is heated in air, and he attributed it to the union of a spirit (gas) with the metal. No notice was taken of this discovery. The increase in weight seems to have surprised and puzzled the chemists in the succeeding centuries. The result was, later on, said to be due to "the absorption of the vapours of charcoal," or, as R. Boyle (1683) expressed it, to "the arresting of igneous corpuscles" which passed through the walls of the vessel in which the metal was calcined, or to "the removal of matter from the calcining vessel." These hypotheses, more or less modified, were in vogue for nearly a century. Jean Rey (1630) appears to have been the first to test the hypotheses by an appeal to experiment.

1. The facts.—In order to clarify the mind, let us review the facts. It would be futile to look for the causes of phenomena before there is agreement as to the facts. Four things are present during the calcination of the metal in air: (1) the containing vessel or crucible; (2) the metal being calcined; (3) the air; and (4) the source of heat. Again the metal and the containing vessel weigh more after the calcination than they did before.

2. The hypotheses.—In applying the inductive method of investigation to these facts, it is necessary to review every rational explanation consistent with the facts, and to examine each hypothesis impartially, since, as emphasized above, it is necessary to show that the explanation finally selected is alone consistent with the facts. This extension of the inductive process might be called the method of exhaustion. It is a mistake to confine the attention to one hypothesis, because that might seriously limit the range of the inquiry. The mind unconsciously assimilates evidence in favour of a pet hypothesis; and a pet hypothesis is apt to grow from a favoured child to a tyrannical master. Four plausible hypotheses may be suggested to explain the cause of the increase in weight: (1) the gases, etc., from the source of heat unite with the containing vessel; (2) the air unites with the containing vessel; (3) the gases from the flame penetrate the crucible, and unite with the metal; and (4) the air unites with the metal.¹

3. Testing the hypotheses by experiment.—By heating the crucible alone, without the metal, no change in weight occurs. This "blank," "dummy," or "control" experiment shows that neither the first nor the second hypothesis will account for the increase in weight of the metal.

¹ The phlogiston theory—that on calcination metals lose a hypothetical substance called *phlogiston*—is discussed later. According to G. E. Stahl (1723), "the fact that metals, when transformed into their calces, increase in weight does not disprove the phlogiston theory, but, on the contrary, confirms it, because phlogiston is lighter than air, and, in combining with substances, strives to lift them, and so decreases their weight; consequently, a substance which has lost phlogiston must be heavier." It may not seem rational to postulate the existence of a substance weighing less than nothing. It will be observed, however, that the assertion: *all matter is heavy and possesses weight*, is only one way of saying that "the attraction of gravitation exists between all masses of matter." This is by no means a self-evident principle, because it is just as easy to conceive of two masses of matter repelling one another, and easier still to conceive of two masses of matter neither attracting nor repelling one another. Hence, the assumption of a phlogiston weighing less than nothing is not so silly as is sometimes supposed. It is quite true that such forms of matter have never been detected, and, accordingly, we assume that they do not exist. Hence also arises the definition of matter indicated in a later chapter.

The third hypothesis can be tested by heating the crucible and the metal out of contact with the air. There is then no change in the weight of the metal. The third hypothesis is therefore untenable. This method was not practicable for the early chemists, and hence Rey employed a less decisive test. It might be expected that if the results depend upon the absorption of the flame gases, different results must be obtained by using different sources of heat—sun-glass, etc.—but the same results are obtained in every case, and accordingly the third hypothesis is probably wrong.

4. The conclusion.—Apparently the only constant factor is air. *The sole invariable antecedent of a phenomenon is probably its cause.* Hence, unless something has been overlooked, we conclude that when metals are heated in air, the increase in weight is due to the fixation of air by the metal, and not to the absorption of furnace gases, nor to variations in the weight of the vessel in which the calcination is made.

Rey also made the interesting unforeseen observation that “nature, in her inscrutable wisdom, has set limits which she does not overstep”; in other words, *however long a metal may be heated in air, a definite weight of each metal can combine with only a definite maximum amount of air.* Students to-day regularly repeat Rey’s experiments on the metals—under various guises—as class exercises. The following table is taken from students’ laboratory notebooks:—

TABLE I.—ACTION OF AIR ON THE CALCINATION OF THE METALS.

Metal.	Weight of metal. Gram.	Weight of calx. Gram.	Increase in weight. Gram.	Ratio. Weight air absorbed to Metal used.
Magnesium	1	1.658	0.658	1 : 1.52
Zinc	1	1.246	0.246	1 : 4.06
Aluminium	1	1.890	0.890	1 : 1.12
Copper	1	1.252	0.252	1 : 3.97
Tin	1	1.269	0.269	1 : 3.72

Hence, one gram of the absorbed air is respectively equivalent to

(Absorbed air).	Magnesium.	Zinc.	Aluminium.	Copper.	Tin.
1	1.52	4.06	1.12	3.97	3.72 grms.

5. Anticipation of new phenomena.—A good hypothesis ought to predict phenomena which have not been observed, and to foretell the results of new experiments; because, if the hypothesis be true, it ought to include all other cases. A hypothesis which is not illogical and which does not contradict known facts is to be judged by its usefulness. The end justifies the means.¹ When the consequences of a hypothesis are logically deduced, a good hypothesis should not only explain, but it should also anticipate facts. Rey’s hypothesis can be used to predict new results. In 1770, A. L. Lavoisier wrote:

Thus did I at the beginning reason with myself. . . . If the increase in weight of a metal calx (calcined in a closed vessel) be not due to the addition of

¹ G. J. Stoney expressed the idea neatly: “A theory is a supposition we hope to be true; a hypothesis is a supposition which we expect to be useful.”

fire matter, nor of any other extraneous matter, but to the fixation of a portion of the air contained in the vessel, the whole vessel, after calcination, must be no heavier than before, and must merely be partly void of air, and the increase in the weight of the vessel will not occur until after the air required has entered.

Lavoisier confirmed this inference experimentally on November 12, 1774; although the gifted Russian chemist, M. W. Lomonosoff, had come to the same conclusion 1756, eighteen years before Lavoisier.

§ 9. Lavoisier's Experiments on the Composition of Air.

Nature is ever making signs to us, she is ever whispering to us the beginnings of her secrets; the scientific man must be ever on the watch, ready at once to lay hold of nature's hint, however small; to listen to her whisper, however low.—M. FOSTER.

Antoine Laurent Lavoisier (1774) extended Rey's experiments with more decisive results. Lavoisier heated tin along with air in a *closed*

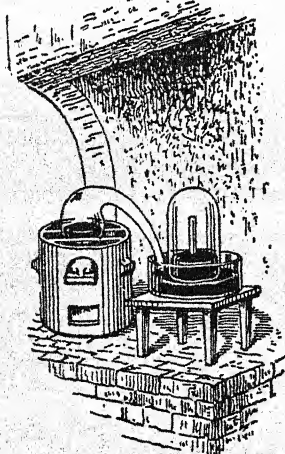


FIG. 2.—Lavoisier's Experiment on the Composition of Air.

vessel. The vessel containing the air and tin did not increase in weight, although part of the air was absorbed. When the flask was opened, air rushed in, and the increase in the weight of the vessel was found to be equal to the increase in weight which the tin alone had suffered. Hence, Lavoisier concluded, with Rey, that the increase in the weight of the tin was due solely to an absorption of the air in which the calcination had occurred. There was not sufficient air in the flask to "saturate" all the tin, and yet some air always remained as a residue. Hence, Lavoisier concluded further that only part of the air can combine with the metal during the calcination; he also found that *the increase in the weight of the tin during calcination is equal to the decrease in the weight of the air*. Hence, it seems as if air contains at least two constituents, only one of which is absorbed by the heated metal. This inference must be tested by experiment.

Lavoisier continued this important work with mercury instead of tin. The mercury was confined in a retort with an S-shaped neck which dipped under a bell-jar in a trough of mercury, as illustrated in Fig. 2. The air in the retort was in communication with the air in the bell-jar. The level of the mercury in the bell-jar was adjusted at a convenient level, and its position "very carefully marked with a strip of gummed paper." By means of a charcoal furnace, the mercury in the retort was heated—not quite to its boiling point. Lavoisier said: "Nothing of note occurred during the first day. The second day I saw little red particles¹ swimming over the surface of the mercury, and these increased in number and volume during four or five days; they then stopped increasing and remained in the same condition. At the expiration of twelve days, seeing that the

¹ The calx or oxide of mercury is red. It is now called "red oxide of mercury," or "mercuric oxide."

calcination of the mercury made no further progress, I put the fire out." After making allowance for variations of temperature and pressure, Lavoisier noticed that the volume of air in contact with the mercury was about 50 cubic "inches," and after the experiment, between 42 and 43 cubic "inches." About one-sixth of the volume of air in the apparatus was absorbed by the mercury.¹ The air which remained in the retort was not absorbed by the hot mercury; it extinguished the flame of a burning candle immersed in the gas;² and a mouse was quickly suffocated when placed in the gas. Hence, Lavoisier called the gas azote, "from the α privative of the Greeks, and $\zeta\omega\eta$, life." In France the gas is still "azote," though in Britain it is called "nitrogen," and in Germany "Stickstoff," that is, "suffocating stuff."

By collecting the red powder and reheating it in a suitable retort, Lavoisier obtained between 7 and 8 cubic "inches" of a gas which had obviously been previously removed from the air by the hot mercury. When a burning candle was immersed in the gas, the candle burnt with "blinding brilliancy," as Lavoisier expressed it; a smouldering splinter of wood burst into flame when plunged in the gas; and the gas did not suffocate a mouse like azote. Lavoisier first called this gas *vital air*, and afterwards *oxygen*. The latter term is its present-day designation. In this manner, Lavoisier proved that atmospheric air is made up of two gases—oxygen and nitrogen—of different and even opposite natures,³ the oxygen alone combines with the metal during calcination.

Assuming that this interpretation of the experiments is correct, Lavoisier inferred that by mixing azote and oxygen in the right proportions, it ought to be possible to reproduce atmospheric air. This Lavoisier did, and the mixture was found to behave with respect to "combustion, respiration, and the calcination of metals similar in every respect to atmospheric air."

Joseph Black (1778) showed that atmospheric air contained a small quantity of what he called "fixed air," which was absorbed by lime-water. Black's "fixed air" is now called "carbon dioxide." In 1777, Carl Wilhelm Scheele inferred that air also contains a little ammonia, because a bottle of hydrochloric acid, when exposed to the air, becomes covered with a deposit of sal-ammoniac (ammonium chloride). So far as Scheele could tell, the only source of the ammonia was the atmospheric air.

Summary.—Summarizing in modern language these results:

1. Atmospheric air is largely made up of oxygen and nitrogen, roughly in the proportion 1 : 4 by volume.

¹ More exact experiments show that one-fifth would be nearer the mark.

² The old chemists used the term *air* where we use *gas* in the sense of an æriform elastic fluid. John Baptista van Helmont (*Oriatrike, or Physick Refined*, posthumous Eng. trans. London, 1662), introduced the word *gas* into chemistry in order to distinguish the vapour given off by water at ordinary temperatures from steam. He said the term "*gas*" is not far severed from the *chaos* of the ancients" (ancients). Just as the "*chaos* of the ancients" was a confused mixture of elements from which the Creator produced the universe, so, to Van Helmont, the vapour of water was a confused mass of elements from which all material substances could be produced. It is an easy transition from *chaos* to *chas*, which has the sound of *gas*—G. F. Rodwell. Some derive the word *gas* from the *geist* (spirit) of the Germans.

³ It must be added that C. W. Scheele (1777) was about the first to state definitely that air "must be made up of two kinds of elastic fluids" (gases).

2. Atmospheric air also contains small traces of ammonia and carbon dioxide.
3. When some metals are calcined in air, they combine with the oxygen of the air to form metallic oxides (calces).

Here then the student has a rich heritage ; he has received in one lesson the result of nearly a thousand years of thought and labour on the nature of atmospheric air !

Difficulties which confronted the early investigation of air.—It seems curious that such a long period of time should have been required to work from Geber's note to the effect that "metals increase in weight when calcined in air," to Lavoisier's proof—1774—that "the increase in weight is due to the absorption of oxygen from the air." This will occasion no surprise when we remember the difference between the properties of air which cannot be seen, and the properties of solids and liquids which can be readily seen and handled. The most obvious property of matter is its visibility, and the conception of matter divested of this quality is no small effort to a mind untutored in physical thought. As G. F. Rodwell has pointed out, the inquiry into the nature of an intangible, invisible body, which exercises no apparent effect on surrounding objects, belongs to an advanced order of experimental philosophy. There were no means of recognizing even the more salient properties of air at the disposal of the chemists until a comparatively late period, and the earlier chemists, accordingly, believed air to be intrinsically different in its essence from more familiar visible substances. To illustrate the ideas about air which prevailed at the end of the 18th century, the opening words of Lavoisier's essay, "On the nature of the substance which combines with metals during their calcination, and which increases their weight" (1774), may be quoted :

Do different kinds of air exist ? Is it enough that a body should be permanently expanded for it to be considered a particular kind of air ? Are the different airs found in nature or formed by us, specific substances, or are they modifications of atmospheric air ?

It will not be always expedient to follow the history of each hypothesis and each conquest of truth, step by step, as in the case of air. That, of course, would be an ideal plan. Goethe was quite right : "The history of a science is the science itself." Unfortunately, we cannot always spare time to wander with the original investigators into the byways of knowledge. This text-book must, therefore, take the place of one of the ancient genii, and jog the student's elbow when he wanders from the right, or rather the conventional, track. We must learn to profit by the experience of others ; and, if possible, leave behind a record of our own so that future generations may profit by our successes and failures.

§ 10. What is an Element ?

I must not look upon any body as a true principle or element, which is not perfectly homogeneous, but is further resolvable into any number of distinct substances.—R. BOYLE.

We have just seen that air can be resolved into two gases—oxygen and nitrogen. It is further possible to resolve all known substances—air, water, etc.—into about eighty distinct, elemental, or primitive forms of

matter. Sulphur and iron, for example, can be obtained from ferrous sulphide; sulphur, iron, and oxygen can be obtained from ferrous sulphate; mercury and oxygen, from mercuric oxide; hydrogen and chlorine, from hydrogen chloride; etc. No chemist, however, has ever separated from oxygen anything but oxygen; from nitrogen, anything but nitrogen; nor from mercury, anything but mercury. Hence, oxygen, nitrogen, and mercury are said to be elements. We can combine mercury, oxygen, and nitrogen in many different ways, and then decompose the resulting compound; but we can get nothing more from the compound than the three elements—mercury, nitrogen, and oxygen—used at the commencement. But we are not yet prepared with a definition of the term "element."

We are indebted to Robert Boyle (1678) and to A. L. Lavoisier (1789) for the modern conception of the word. Lavoisier, quite logically, considered lime, magnesia, baryta, and alumina to be elements. We now know that these "elements" of Lavoisier are compounds of oxygen with calcium, magnesium, barium, and aluminium respectively. Lavoisier apparently foresaw some such possibility, for he stated: "We are certainly authorized to consider them simple bodies until by new discoveries, their constituent elements have been ascertained." Again, in 1811, the question whether chlorine—then called oxymuriatic gas—was really an element or a compound of oxygen with some other element was raised by Humphry Davy. Davy claimed that chlorine is an element because, although oxygen was believed to be present, none could be found. "Hence," added Davy, "we have no more right to say that oxymuriatic gas (*i.e.* chlorine) contains oxygen than to say that tin contains hydrogen. . . . Until a body is decomposed, it should be considered simple."

It is not possible to improve upon Lavoisier's conception of an element, and I feel compelled to quote his words, although written before 1789:

If we apply the term elements or principles to bodies to express our idea of the last point which analysis is capable of reaching, we must admit, as elements, all substances into which we are able to reduce bodies by decomposition. Not that we are entitled to affirm that these substances which we consider as simple, may not themselves be compounded of two, or even of a greater number of more simple principles; but since these principles cannot be separated, or rather, since we have not hitherto discovered the means of separating them, they are, with regard to us, as simple substances, and we ought never to suppose them compounded until experiment and observation have proved them to be so.

In fine, "element" is a conventional term employed to represent the limit of present-day methods of analysis or decomposition. We may therefore summarize these ideas in the definition: An element is a substance which, so far as we know, contains only one kind of matter. To say that the elements cannot be decomposed is an unwarranted reflection on the powers of our successors. The moment Auer von Welsbach (1885) proved that didymium was a mixture of praseodymium and neodymium, one element ceased to exist, and two elements were born. If it were found to-morrow that the element chlorine is really a compound of two new elements previously unknown, the fact would be important and it would change the face of chemistry, but it would not render useless any facts we know about chlorine.

The old alchemists sought to transform some of the common metals into gold. Whenever the attempt has been made with materials known

to be free from gold, no transmutation has been observed. There is nothing intrinsically absurd in the notion, but at present, no authentic transmutation of this kind has been deliberately, or rather intentionally, accomplished. When the evidence has permitted a critical examination, every recorded instance has been traced to a mal-observation¹; and evidence which cannot be tested is outside the range of scientific methods.

§ 11. The Four-Element Theory.

Four elements intimately united
Form the whole world.—SCHILLER.

The four-element theory is one of the oldest attempts to classify the multitudinous forms of matter which make up the world. To the early philosophers, the world was composed of four distinct principles or entities—the *earth* typified all solids; *water*, liquids; *air*, the winds, clouds, and the breath; and lastly *fire* which was the symbol of the sun, and worshipped by many as a god. Hence, in the writings of the alchemists, we usually find a chapter devoted to this quartet: earth, water, air, and fire. The early philosophers added a fifth element—*quinta essentia*—perhaps analogous with the primal matter of the Greeks. This was supposed to be a subtle extract, the quintessence of the other four. The ancient Hindu philosophers had previously added a fifth element which, in their system, was supposed to be the medium for propagating sound, etc., and which, in consequence, had something in common with the modern concept of an æther pervading all space.²

While accepting the hypothesis of a universal primal matter, many of the alchemists of the Middle Ages rejected the four-element theory and adopted three principles—*tria prima*—which they called respectively salt, sulphur, and mercury. *Salt* represented the *principle of fixity or solidity*; *mercury* typified air and water, or the *principle of liquidity and gaseity*; and *sulphur* typified fire, or the *principle of combustion*. These three principles were thus parallel with the four elements of the Greeks, and it is now difficult to see what was gained by the change. Possibly they had some idea of making a mundane trinity recall The Trinity.

The three- and four-element theories were demolished when water, air, and the earths were decomposed into still simpler bodies; and when fire was shown to be a manifestation of energy. It is probable that the term "element" was not used by the old philosophers in the same sense that it is to-day. Whatever the idea involved in the four-element theory, it was believed by many different races in different parts of the globe; it has pervaded the philosophy of all thinking races; it has been sung by the poets of every land; and it has had a longer life than any succeeding philosophy. The theory was living a couple of centuries ago; it is now dead.

¹ See the chapter on "Radioactivity" towards the end of this work.

² Air, not the æther, is the medium for propagating sound.

CHAPTER II

COMBINATION BY WEIGHT

§ 1. The Law of Constant Composition.

Nature in her inscrutable wisdom has set limits which she never oversteps.—JEAN REY.

ATTENTION must now be directed to the singular observation made by Jean Rey (1630) that during the calcination of a metal in air, "the weight of the metal increased from the beginning to the end, but when the metal is saturated, it can take up no more air. Do not continue the calcination in this hope: you would lose your labour." The examples previously quoted—Table I.—have shown that one gram, and only one gram, of air is absorbed by definite amounts of the given metals under the conditions of the experiment, and Lavoisier's work proves that the *oxygen* of the air is alone absorbed. Accordingly, one part by weight of oxygen is equivalent to:

Oxygen.	Magnesium.	Zinc.	Aluminium.	Copper.	Tin.
1	1.52	4.06	1.12	3.97	3.72

Instead of taking the weight of oxygen unity, it will be more convenient, later on, and also more in accord with general usage, to make oxygen 8 instead of unity. Hence, multiplying the preceding numbers by 8, we obtain:

Oxygen.	Magnesium.	Zinc.	Aluminium.	Copper.	Tin.
8	12.16	32.48	8.96	31.76	29.76

When magnesium is calcined in the presence of oxygen, or air, the metal always unites with the oxygen in the proportion of one part of oxygen per 1.52 parts of magnesium, or 8 parts by weight of oxygen per 12.16 parts by weight of magnesium. The same principle obtains when magnesium oxide is made in several different ways; and likewise with the other metallic oxides. Hence, as P. G. Hartog puts it: two like portions of matter have the same composition. The converse of this statement is not necessarily true.

The exact work of J. S. Stas and of T. W. Richards and many others has firmly established this deduction for the regular type of chemical compounds. J. S. Stas (1860), for example, studied among other things, the composition of silver chloride prepared by four different processes at different temperatures. He found that 100 parts of silver furnished 132.8425, 132.8475, 132.842, 132.848 parts of silver chloride; and that neither the temperature nor the method of preparation had any influence

on the composition of the chloride. The difference between the two extremes is less than 0.006 part per 100 parts of silver. This shows that the errors, incidental to all experimental work, are here remarkably small. Hence, Stas stated: "If the recognized constancy of stable chemical compounds needed further demonstration, I consider the almost absolute identity of my results has now completely proved it."

The student will take notice that *we are unable to prove the law of constant proportions with mathematical exactness. However skilful a chemist may be, it is impossible to make an exact measurement without committing an "error of observation" or an "error of experiment."* It is assumed that the small difference 0.005 per cent. between the two extreme results of Stas (1) is wholly due to the unavoidable errors of experiment, for we cannot expect an exact solution of the problem; and (2) is not due to a very slight inexactitude in the law of constant proportions. (Cf. p. 570.)

The composition of a definite chemical compound appears to be independent of its mode of formation, and therefore it is inferred that substances always combine in definite proportions. If an excess of one substance be present, the amount in excess is extraneous matter. This deduction from the observed facts is called the law of definite proportions, or the law of constant composition: a particular chemical compound always contains the same elements united together in the same proportions. Probably no generalization in chemistry is more firmly established than this. It was not discovered by any particular man, but gradually grew among the doctrines of chemistry. The law was tacitly accepted by many before it was overtly enunciated—e.g. J. Rey (1630), I. Newton (1706), G. E. Stahl (1720), F. G. Rouelle (1764), C. F. Wenzel (1777), T. Bergman (1783), etc. So great is the faith of chemists in the truth of this generalization that a few accurate and careful experiments are considered sufficient to settle, once for all, the composition of a substance. For instance, if a substance possessing all the properties of magnesium oxide be given to a chemist, without taking any more trouble, he knows that it will contain 12.16 parts of magnesium for every 8 parts of oxygen.

Historical.—The validity of the law was the subject of an interesting controversy during the years between 1800 and 1808. J. L. Proust maintained that constant composition is the invariable rule; C. L. Berthollet maintained that constant composition is the exception, variable composition the rule. Proust's words are worth quoting:

According to my view, a compound is a privileged product to which nature has assigned a fixed composition. Nature never produces a compound, even through the agency of man, other than balance in hand, *pondere et mensura*. Between pole and pole compounds are identical in composition. Their appearance may vary owing to their manner of aggregation, but their properties never. No differences have yet been observed between the oxides of iron from the South, and those from the North; the cinnabar of Japan has the same composition as the cinnabar of Spain; silver chloride is identically the same whether obtained from Peru or from Siberia; in all the world there is but one sodium chloride; one saltpetre; one calcium sulphate; and one barium sulphate. Analysis confirms these facts at every step.

It might be thought that positive assertions of this kind, backed by accurate experimental work, would leave no subject for disputation. But, surveying the battlefield in the light of the present-day knowledge, it

seems that another quite different phenomenon was confused with the law of constant composition; and the methods of analysis were not very precise. Some, probably from the unfounded belief that "Proust deservedly annihilated Berthollet," call the generalization discussed in this chapter, "Proust's law." We shall see later that a phenomenon which Proust apparently did not clearly recognize prevented him from annihilating Berthollet.

§ 2. Physical and Chemical Changes.

Most of the substances belonging to our globe are constantly undergoing alterations in sensible qualities, and one variety becomes, as it were, transmuted into another. Such changes, whether natural or artificial, whether slowly or rapidly performed, are called chemical. Thus, the gradual and almost imperceptible decay of the leaves and branches of a fallen tree exposed to the atmosphere and the rapid combustion of wood in our fires, are both chemical operations.—H. DAVY.

One element or compound is distinguished from all other elements or compounds in possessing certain specific and characteristic properties; or, in the words of an old alchemist, "God hath sealed each substance with a particular idea." First and foremost, a chemical compound has a fixed and definite composition; then again, it melts and boils at definite temperatures; its crystalline form, specific gravity, specific heat, colour, odour, behaviour when in contact with other substances, etc., are characteristic of one particular chemical compound. When the melting point of, say, pure silver chloride has been once accurately determined, it follows that all other samples of pure silver chloride will melt at the same temperature under the same conditions. The more salient characteristic properties of an element or compound are employed as tests for its identification—that is, for distinguishing it from all other known compounds. Thus a student would be probably correct in stating that a solution contained a silver compound if it gave a white precipitate when acidified with hydrochloric acid, and the precipitate was insoluble in hot water, and soluble in aqueous ammonia.

Physical changes.—When liquid water becomes ice or steam there is no change in the *chemical* nature of the substance, for the matter which makes steam and ice is the same in kind as that of liquid water. A substance can generally change its state, as when liquid water becomes steam or ice. The idea is further emphasized by the fact that in most cases a substance is called by the same name, whether it be in the solid, liquid, or gaseous state of aggregation, *e.g.* we speak of "liquid" oxygen, "liquid" air, "molten" silver chloride, etc. Again, matter may change its *volume* by expansion or contraction; it may change its *texture*, as when a porous solid is compressed to a compact mass; it may change its *form*, as when matter in bulk is ground to powder; it may change its *magnetic qualities*, as when a piece of soft iron in contact with a magnet attracts other pieces of iron, etc. It is conventionally¹ agreed to say that in none of these cases

¹ I must confess that in writing this book I have found this chapter to be the most difficult. We have some uncomfortable doubts if magnetized and demagnetized iron can be called the same kind of matter; similar remarks apply to, say, monoclinic and rhombic sulphur; and to water at -20° , $+20^{\circ}$, and $+120^{\circ}$. The student will appreciate the difficulty after reading § 3 in the chapter on "Water," and § 7 in the chapter on "Sulphur."

of physical change is there any evidence of the formation of a new substance; and that the matter does not lose or change those properties which distinguish it from other forms of matter. A physical change involves an alteration in the properties of a substance without the formation of a new substance.

Chemical changes.—When magnesium metal is heated in air, a white powder is formed, and when mercuric oxide is similarly treated, mercury and oxygen are obtained. The action of heat in both cases furnishes forms of matter with very different specific properties from those forms of matter employed at the start. A chemical change involves the formation of a fresh substance or substances, with different specific properties from the original substance or substances. In both chemical and physical changes, as we shall soon find, the total mass of matter before and after the change remains constant, but in chemical changes alone the *kind* of matter alters.

It is not always easy to distinguish between physical and chemical changes, because the only real distinction between the two turns on the question: Is there any evidence of the formation of a new substance during the change? The evidence, as we shall soon see, is not always conclusive.

§ 3. Compounds and Mixtures.

The common operations of chemistry give rise in almost every instance to products which bear no resemblance to the materials employed. Nothing can be so false as to expect that the qualities of the elements shall be still discoverable in an unaltered form in the compound.—W. WHEWELL.

1. The constituents of a compound are combined in definite proportions.—The law of constant proportions is of fundamental importance in forming a conception of the meaning of the term "chemical compound." If a substance produced in different ways be not constant in composition, it is not considered to be a chemical compound, but rather a mixture. R. Bunsen (1846), for example, showed that the proportion of oxygen to nitrogen in atmospheric air is not constant, because the oxygen varies from 20.97 to 20.84 per cent. by volume, by methods of measurement with an error not exceeding 0.03 per cent. Hence, the oxygen and nitrogen in atmospheric air are said to be simply mixed together, and not combined chemically. We shall soon see, however, that substances with a definite composition are usually, but not always, chemical compounds.

2. Compounds are homogeneous, mixtures are usually heterogeneous.—It is comparatively easy to detect particles of sugar and sand in a mixture of the two; and a simple inspection of a piece of Cornish granite will show that it is a mixture of at least four constituents—silvery flakes of mica; black patches of schörl; whitish crystals of feldspar; and clear glassy crystals of quartz. A photograph of a thin slice of this rock, as it appears under the microscope magnified about 50 diameters, is shown in Fig. 3. Although the particles of feldspar, mica, schörl, and quartz differ from one another in size and shape, no essential difference can be detected in the composition and properties of different samples of pure quartz, pure feldspar, mica, and schörl. Hence, it is inferred that the sample of granite is a mixture of schörl, feldspar, quartz, and

mica; and that each of these minerals is a true chemical compound. Very frequently, the constituents of a mixture are too small to be distinguished by simple inspection, and the body appears homogeneous. A microscopic examination may reveal the heterogeneous character of the substance. Blood and milk, for instance, appear to be homogeneous fluids, but under the microscope the former appears as a colourless fluid with red corpuscles in suspension; and milk appears as a transparent liquid containing innumerable white globules (fat). Naturally, too, the stronger the magnification, the greater the probability of detecting whether the body is homogeneous or not. Sometimes the microscope fails to detect non-homogeneity under conditions where other tests indicate heterogeneity.¹

Before constant composition can be accepted as a proof of chemical combination, it must also be shown that the substance is homogeneous. A

homogeneous substance is one in which every part of the substance has exactly the same composition and properties as every other part. A substance may have a fixed and constant composition and yet not be homogeneous—*e.g.* cryohydrates and eutectic mixtures to be described later. A substance may be homogeneous, for all we can tell to the contrary, and yet not have a constant composition—*e.g.* atmospheric air; a solution of sugar in water, etc. This simply means that all chemical



FIG. 3.—Cornish Granite ($\times 50$).

compounds are homogeneous, but all homogeneous substances are not chemical compounds. Indeed, it is sometimes quite impossible to tell by any single test whether a given substance is a mixture or a true chemical compound.

3. The constituents of a mixture can usually be separated by mechanical processes.—The properties of a mixture of finely powdered iron and sulphur have been used in chemical text-books since 1823 to illustrate the difference between mixtures and compounds. It would be difficult to find a better example. Rub together a mixture containing, say, 6 grams of iron and 4 grams of sulphur in a mortar, and note that: (1) the colour of the mixture is intermediate between the colour of the iron and of the sulphur; (2) the particles of iron and sulphur can be readily distinguished under the microscope; (3) some of the iron can be removed without difficulty by means of a magnet; and (4) the two can be separated

¹ See a later section on "Ultramicroscopic Particles." It may seem curious to refer a student to a later chapter. The reference, of course, is intended when the book is read a second time, not the first time. A text-book should be read forwards and backwards.

quite readily by washing the mixture on a dry filter paper by means of carbon disulphide. The sulphur dissolves in the carbon disulphide; the solution can be collected in a dish placed below the filter paper; and the sulphur can be recovered by allowing the carbon disulphide to evaporate from the dish. Sulphur remains behind as a crystalline residue (Fig. 162). The metallic iron remains on the filter paper. Here then the constituents of the mixture have been separated by the mechanical processes—magnetizing, and the action of solvents. It is not always possible to apply these tests. Solvents, as we shall find later, sometimes decompose a compound into its constituents, or conversely, “cause” the constituents of a mixture to combine.

4. A mixture usually possesses the common specific properties of its constituents; the properties of a compound are usually characteristic of itself alone.—The properties of a mixture are nearly always additive, *i.e.* the resultant of the properties of the constituents of the mixture. For instance, a mixture of equal parts of a white and black powder will be grey. The specific gravity of a mixture of equal volumes of two substances of specific gravity ¹ 3 and 5 will be 4, because if 1 c.c. of water weighs 1 gram, there will be a mixture of 0.5 c.c. weighing 1.5 gram of one substance; 0.5 c.c. of the other substance weighing 2.5 grams; and $1.5 + 2.5 = 4$ grams per c.c. It must be added that such properties of compounds are additive, for they are the sum of the properties of their constituents.

EXAMPLES.—(1) What is the specific gravity of air containing a mixture of one volume of oxygen and four volumes of nitrogen when the specific gravity of oxygen is 16, and the specific gravity of nitrogen 14.01? One-fifth volume of oxygen weighs 3.2 units, and four-fifths volume of nitrogen weighs 11.2 units. Hence, one volume of the mixture will weigh 14.4 units when one volume of oxygen weighs 16 units.

(2) Ozonized air—a mixture of air and ozone—has a specific gravity 1.3698, and it contains 13.84 per cent. by weight of air, specific gravity unity, and 86.16 per cent. of ozone. What is the specific gravity of ozone? Here 13.84 grams of air occupy $13.84 \div 1$ volume; and 86.16 grams of ozone occupy $86.16 \div x$ volumes, where x denotes the specific gravity of ozone. Hence, 100 grams of ozonized air occupy $100 \div 1.3698 = 73$ volumes. Hence, $73.00 = 86.16 \div x + 13.84$; or $x = 1.456$.

If a portion of the mixture of sulphur and iron indicated above be placed in a hard glass test-tube, and warmed over the Bunsen's flame, the contents of the tube begin to glow and a kind of combustion spreads throughout the whole mass. When cold, break the test-tube, and note that (1) the porous black mass formed during the action is quite different from the original mixture; (2) the microscope shows that the powdered mass is homogeneous; (3) it is not magnetic like iron,² and (4) it gives

¹ SPECIFIC GRAVITY.—The student is supposed to know that *specific gravity is a number which expresses how much heavier a given substance is than an equal volume of water taken at a standard temperature*. In the case of gases, it may be that air = unity, oxygen = 16, hydrogen = 1, or hydrogen = 2 is taken as standard; and in the case of liquids and solids, water at + 4°, or at 0°, is taken as unity. The great value of specific gravity data lies in the fact that *specific gravity is a number which enables volume measurements to be converted into weights, and weight measurements to be converted into volumes*. Specific gravity may thus be regarded as the weight of unit volume if the standard water = 1 be taken, and the weights are reckoned in grams, and volumes in cubic centimetres. There is no need here to distinguish between density and specific gravity.

² This provided the iron was not in excess.

up no sulphur when digested with carbon disulphide.¹ These facts lead to the assumption that there has been a chemical reaction between the sulphur and the iron. *When chemical combination occurs, the reacting constituents appear to lose their individuality or identity more or less completely, and each new substance which is formed has its own distinctive properties.*

5. Thermal, actinic (light), or electrical phenomena usually occur during chemical changes.—Attention must be directed to the fact that a great deal of heat was developed during the combination of the iron and sulphur. The heat required to start the reaction does not account for the amount of heat developed during the reaction. This point is perhaps better emphasized by placing an intimate mixture of powdered sulphur and zinc on a stone slab. After the flame of a Bunsen's burner has been allowed to play on a portion of the mixture for a short time to start the reaction, the zinc and sulphur combine with almost explosive violence. Large amounts of heat and light are developed during the reaction.

If a plate of commercial zinc be placed in dilute sulphuric acid, bubbles of gas are copiously evolved, and if a thermometer be placed in the vessel, the rise of temperature shows that heat is generated during the chemical action. If the zinc be pure, very little if any gas

is developed. It makes no difference if a plate of platinum be dipped in the same vessel as the zinc, provided the plates are not allowed to come into contact with one another. If the two plates are connected by a piece of copper wire, a rapid stream of gas bubbles arises from the surface of the platinum plate, and some gas also comes from the zinc plate. The platinum is not attacked by the acid in any way, but the zinc is rapidly dissolved. If a voltmeter be interposed in the circuit between the two plates—Fig. 4—the deflection of the needle shows that an electric current “passes” from the platinum to the zinc, as represented by the arrows. The electric current is generated by the chemical reaction between the zinc and the acid, which results in the formation of zinc sulphate and a gas. The action will continue until all the acid or the zinc is used up.

Nomenclature.—The junction of the wire with the zinc plate is conventionally called the negative or — pole; and the junction of wire with the platinum plate is called the positive or + pole. For convenience, the zinc or positive plate

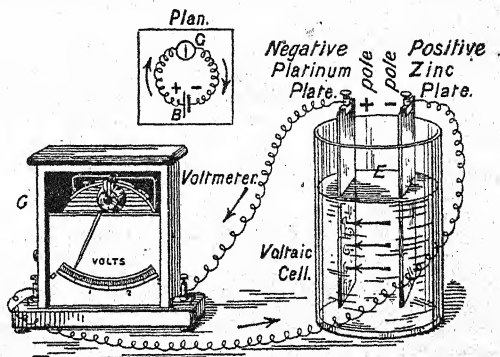


FIG. 4.—Electricity by Chemical Action—Voltaic Cell.

¹ This provided the sulphur was not in excess, but the excess of sulphur if present can often be driven off as vapour.

of the cell *B* is often represented by a short thick line, and the platinum or negative plate by a longer thinner line as illustrated by the plan, Fig. 4. Hence the platinum in this simple cell is called the positive pole, or negative plate; and the zinc is the negative pole, or positive plate. Here *G* represents the voltmeter or galvanometer. The vessel of acid with its two plates is called a voltaic cell, and this particular combination can be symbolized:—

Platinum | Dilute sulphuric acid | Zinc

The chemical reaction just indicated is far from being the most economical mode of generating electricity, but all the different forms of voltaic cell on the market agree in this: Electricity is generated during chemical action.

The developments of heat, light, or electrification are common concomitants of chemical action. The absence of such phenomena when substances are simply mixed together is usually taken as one sign that chemical action has not taken place. When nitrogen and oxygen are mixed together in suitable proportions to make atmospheric air there is no sign of chemical action, and this fact is sometimes cited among the proofs that air is a mixture. The argument is not conclusive, because the condensation of water vapour and the freezing of liquid water are often cited as examples of physical change although heat is evolved in both transformations.

Summary.—The tests for distinguishing chemical compounds from mixtures involve answers to the following questions:

1. Are the different constituents united in definite and constant proportions?
2. Is the substance homogeneous?
3. Are the properties of the substance additive?
4. Were thermal, actinic, or electrical phenomena developed when the substance was compounded?
5. Can the constituents be separated by mechanical processes?

Mechanical processes of separation.—The so-called mechanical processes of separation usually include: (1) Magnetizing, hand-picking, sieving, etc.; (2) If some mixtures be placed in liquids of the right specific gravity, the lighter constituents will float and the heavier constituents will sink; (3) Differences in the solubility of the constituents in suitable solvents; (4) Distillation; (5) Freezing; (6) Liquation; (7) Diffusion; (8) Elutriation; (9) Flotation, etc.

It may be useful to again emphasize the fact that the so-called "mechanical" processes of separation, involving solution, freezing, and distillation, are not always satisfactory tests for distinguishing chemical compounds from mechanical mixtures. It is generally stated that "a solution of sugar or of salt in water is a mechanical mixture because, though homogeneous, the salt or sugar can be recovered unchanged from the water by the mechanical process of evaporation." This is an unwarranted assumption. The salt and water may have combined, and the product of the chemical combination may be decomposed into salt and water during the process of evaporation.

The above list does not exhaust the available tests, but in spite of what we know, there is sometimes a lingering doubt whether a particular substance is a mixture or a true chemical compound. This arises from the fact that some of the tests are impracticable, others are indecisive. As previously stated, owing to our ignorance, it is not always easy to state "the truth and nothing but the truth."

§ 4. Circumstantial and Cumulative Evidence.

To find the truth is a matter of luck, the full value of which is only realized when we can prove that what we have found is true. Unfortunately, the certainty of our knowledge is at so low a level that all we can do is to follow along the lines of greatest probability.—J. J. BEZELIUS.

Suppose a substance is suspected to be a chemical compound because it appears to be homogeneous; on investigation, we find that it has a fixed definite composition. This verifies our first suspicion, and the joint testimony gives a very much more probable conclusion than either alone. By piling up the evidence in this manner, for or against our suspicion, we can make a chain of circumstantial evidence which enables highly probable conclusions to be drawn. Each bit of evidence taken by itself is not of much value, but all the evidence taken collectively has tremendous weight. It is easy to see, too, that the probability that an hypothesis is valid becomes less as the number of unproved assumptions on which it is based becomes greater. On the other hand, plausible hypotheses neatly dovetailed may sometimes fit together so well as to strengthen rather than weaken one another; but the truth of the hypotheses is not thereby established.

We can even get a numerical illustration. If the definite-compound test be right nine times out of ten, the probability that a given substance of definite composition is not a true compound is $\frac{1}{10}$; similarly, if the homogeneous test be right three times out of four, the probability that the given homogeneous substance is not a chemical compound is $\frac{1}{4}$; and the probability that the given homogeneous substance of definite composition is not a true compound is $\frac{1}{40}$. Every bit of additional evidence in favour of a conclusion multiplies the probability of its being correct in an emphatic manner; and evidence against a conclusion acts similarly in the converse way. Huxley has stated that one of the tragedies in science is the slaughter of a beautiful hypothesis by one *incongruent fact*; a conclusion based solely upon circumstantial evidence is always in danger of this Damoclean sword.

A writer has said: "When two facts seem to be in conflict, we may be driven to decide which is the more credible of the two." This statement may give rise to a misunderstanding. We cannot admit the possibility of two contradictory facts. Facts can, and often do, contradict hypotheses. Again, a fact is a fact and cannot be disputed. If there be any doubt about the truth of an alleged fact, something is wrong. The laboratory, not the study, is the place to decide if the alleged fact is the result of an incomplete or of a mal-observation. Facts *quâ* facts cannot be graded in degrees of probability or credibility; they cannot be disputed; and they are the last court of appeal.

§ 5. Analysis and Synthesis.

It is surely not fitting for a chemist to make a large number of experiments with the sole object of rapidly making new products, for he will then overlook phenomena and changes during the operations which might serve as important clues to an explanation of nature's secrets.—M. W. LOMONOSOFF.

The term *synthesis*—from the Greek σύν (syn), with; τιθέω (titheo), I place—is employed for the operations involved in making a particular

compound from its constituents. *E.g.* methods for the synthesis of ferrous sulphide, and also for the synthesis of various oxides, were described in preceding sections. The term *analysis*—from the Greek *ἀν* (ana), back; *λύω* (lyo), I loosen—is employed for the process of separating the constituents of a compound or mixture. Thus mercuric oxide is broken down into its constituents when heated. The object of the analysis may be to answer the question: What are the constituents of the mixture or compound? The analysis is then said to be *qualitative*. If the relative quantities of the different constituents are to be determined, the analysis is said to be *quantitative*. For instance, if a weighed portion of a mixture of sulphur and iron be treated with carbon disulphide as described above, and the separated sulphur and iron be weighed, the two weights should be nearly equal to the weight of the original mixture taken for the analysis. The numbers so obtained express the result of a quantitative analysis of the mixture.

Analysis of gunpowder.—Gunpowder is a mixture of nitre (soluble in water), sulphur (soluble in carbon disulphide), and carbon (insoluble in both the solvents just mentioned). Hence, gunpowder can be analysed by first washing a weighed quantity of the powder on a filter paper with warm water and collecting and evaporating the filtered solution to dryness in a weighed dish. The increase in the weight of the dish with its contents represents the amount of nitre. The insoluble residue is dried and treated in a similar manner with carbon disulphide, and the amount of sulphur determined as in the case of the mixture of iron and sulphur. The dried carbon is then weighed. The result of a quantitative analysis of a sample of gunpowder, expressed in percentage numbers, is: nitre, 78 per cent.; sulphur, 12 per cent.; charcoal, 10 per cent. For the appearance of the residual nitre (potassium nitrate) left on evaporating the aqueous solution see Fig. 196 (left); and for the sulphur residue, Fig. 162.

There was one period in the history of chemistry when the discovery or synthesis of new substances was considered to be the main aim of the chemist. New substances were made unmeasured, unclothed, and required to be investigated all over again. The style of some old text-books on chemistry was not far removed from that of cookery-recipe books. Their long monotonous lists of methods of preparing compounds led E. J. Mills to say that "chemistry has become an art of breeding (new compounds)." This work has been useful, for it has furnished modern chemistry with raw empirical material to be worked up into science; indeed, a great deal more empirical data is now available than chemistry has been able to assimilate. Modern chemistry therefore is not so much directed to the discovery of new compounds, as to a more careful study of the old. We are beginning to recognize the truth of the inspired words of M. W. Lomonosoff, cited above, though written in 1751, and the growing use of "squared paper" in chemical text-books is "a sign of the times."

Synthesis of zinc sulphate.—The solution which remains when the dilute sulphuric acid, indicated on p. 33, can dissolve no more zinc may be filtered and evaporated over a hot plate until a drop of the hot solution crystallizes when placed on a cold glass plate. Crystals of zinc sulphate will separate as the solution cools. By evaporating a large volume of the solution very slowly, crystals over a foot long have been obtained. The

appearance of the crystals which separate from a drop of solution slowly evaporated is indicated in Fig. 5, and an outline drawing of a perfect crystal is shown on the same diagram. This experiment illustrates the synthesis of zinc sulphate from metallic zinc and dilute sulphuric acid.

The analysis of aqueous solutions of zinc sulphate by the electric current.—In the experiment illustrated by Fig. 4, an electric current was developed during the reaction between dilute sulphuric acid and metallic zinc which resulted in the formation of zinc sulphate. Fit up a similar arrangement as before, but place two platinum plates, *E*, and pure distilled water in the clean glass jar, which will now be called the "electrolytic cell."

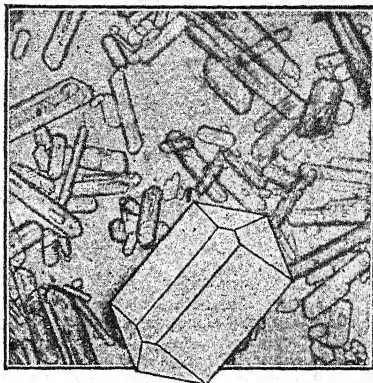


FIG. 5.—Crystals of Zinc Sulphate ($\times 50$).

Connect the two platinum plates with an accumulator or secondary battery, and a voltmeter as indicated in Fig. 6. The object of the accumulator is to generate an electric current.¹ If the water is pure, the needle of the voltmeter moves very little, if at all. Add a concentrated solution of zinc sulphate to the water in the glass jar. The jump of the needle of the voltmeter shows that a current of electricity is flowing through the circuit, and hence also through the solution of zinc sulphate. If chloroform, benzene, or an aqueous solution of cane sugar had been used in place

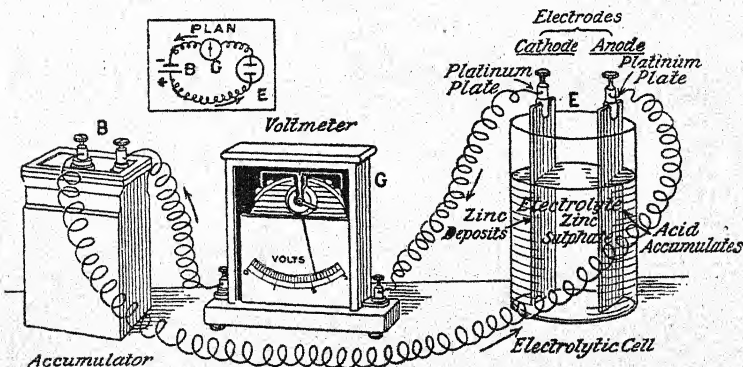


FIG. 6.—Chemical Action induced by Electric Current—Electrolysis.

of the solution of zinc sulphate in the electrolytic cell, no current would pass through the circuit. Hence liquids may be either conductors or

¹ This is a better way of using electricity than working with a primary battery, Fig. 4.

non-conductors of electricity. The current which passes through the solution of zinc sulphate produces some remarkable changes: (1) a spongy mass of metallic zinc accumulates about one of the platinum plates; (2) if the solution be tested, particularly in the neighbourhood of the other platinum plate, sulphuric acid will be found to be accumulating in the solution during the process of electrolysis: and (3) bubbles of oxygen gas, easily tested by collecting some in a test-tube, rise from the same platinum plate about which the acid accumulates. If the experiment be continued long enough, and the products of the action be examined, we shall find that metallic zinc and sulphuric acid have been produced. If the accumulator be disconnected, and the connections be made as indicated in Fig. 4, the zinc will redissolve in the acid, re-producing zinc sulphate, and an electric current will be generated during the dissolution of the zinc.

Nomenclature.—The process of decomposition or analysis by the aid of the electric current is called electrolysis. The liquid which is decomposed is called the electrolyte. The passing of the electric current through the conducting copper wires, and through the conducting platinum plates, produces no change in these metals. Hence, we recognize two kinds of conductivity—in one the conducting medium is decomposed by the current—electrolyte; and in the other the conducting medium is not decomposed by the current—non-electrolyte. The plate at which the zinc collects is called the cathode—from the Greek *κάτα* (*kata*), down; *ὁδός* (*hodos*), a path—and the other plate, about which the acid collects, is called the anode—from the Greek *ἀνά* (*ana*), up; *ὁδός* (*hodos*), a path. The anode and cathode together are called the electrodes. With the conventions indicated on p. 33, as to the direction of the electric current, the current is said to enter the electrolytic cell *via* the anode, and to leave the cell *via* the cathode. The two electrodes are thus “the doors or ways by which the current passes into or out of the decomposing body.” It seems as if the electric current first splits the decomposing liquid into two parts which pass to the electrodes. The term anions—from the Greek *ἰόν* (*ion*), traveller—is applied to those parts of the decomposing fluid which go to the anode; and those passing to the cathode are called cations, and when reference is made to both the anions and cations, the term ions is employed. “Ions” is thus a general term for those bodies which pass to the electrodes during electrolysis. This notation was proposed by M. Faraday in 1834. The subject will be further developed in later chapters.

The principle of reversibility.—The experiments indicated above illustrate an important principle—the principle of reversibility: If an antecedent event A produces an effect B, then an antecedent event B will reproduce the effect A. Thus, chemical action can produce an electric current, Fig. 4, and conversely, an electric current results in a loss of chemical activity (in the battery), Fig. 6. The one can undo the work of the other. Other examples of the principle will be found as we progress in our studies. Heat causes gases to expand; conversely, if a gas expands by its own elastic force, heat will be lost and the gas cooled, etc.

§ 6. Pure Substances.

Pure water is never found in nature. One may even say that no man has ever seen or handled absolutely pure water. It is an ideal substance, to which some specimens of highly purified water have nearly approached.—M. M. P. MUIR.

The substance we call water has its own specific properties, but sea-water, spring-water, rain-water, and distilled water show certain differences

in their properties. The differences, however, are not due to the water, but to the substances—impurities—which the water has dissolved from its surroundings. If sea-water be distilled, the “impurities”—sodium chloride, magnesium chloride, etc.—remain behind. Sea-water is therefore a homogeneous substance, but, rightly or wrongly, it is often stated to be a mixture, because water can be separated by simple evaporation or by freezing. Table salt is more or less impure sodium chloride. The presence of a little magnesium chloride in table salt makes the salt more hygroscopic, so that the contaminated table salt deliquesces ¹ more readily than if magnesium chloride were absent.

Lavoisier's experiments on the transformation of water into earth.—A compound may be contaminated with impurities in many ways—from the raw materials used in preparing the compound; from the vessels in which it was prepared or stored; by exposure to the atmosphere; by the partial decomposition of the substance when exposed to light, etc. It was once believed that air can be condensed to water, as was thought to be proved by “falling dew”; and that water can be changed into an earth, as is evidenced by the residue obtained when rain-water or distilled water is evaporated to dryness in glass vessels. Lavoisier read a paper in 1770 with the idea of “settling by decisive experiments whether water can be changed into earth as was thought by the old philosophers, and still is thought by some chemists of the day.” By distilling water in hermetically sealed glass vessels, weighed before and after the experiment, it can be proved: (1) The earth does not come from outside the vessel, because the weight of the vessel *and* its contents does not alter; (2) The earth does not come from the water, because the weight of the water remains the same before and after the experiment; (3) The earth comes from the vessel, because the vessel loses in weight; and (4) The earth comes *wholly* from the vessel, because the loss in weight of the vessel is virtually equal to the weight of the earth formed. Hence, adds Lavoisier, “it follows from these experiments that the greater part, possibly the whole of the earth separated from rain-water by evaporation, is due to the solution of the vessels in which the water has been collected and evaporated.”

The purity of commercial compounds.—The term “*pure*,” or “*chemically pure*,” is unfortunately used when it is desired to emphasize the fact that a substance is not contaminated with sufficient impurity to appreciably influence the most exact work for which the substance is to be employed. There cannot be degrees of “purity.” A thing is either “pure” or “impure.” It may be convenient to use terms like “highly pure,” “all but pure,” “very impure,” etc., but the term “chemically pure” in the sense of “nearly pure” is objectionable. “Chemically pure” substances, paradoxical as it may seem, are sold with a statement on the labels indicating what impurities are present as well as how much of each. A *commercial* reagent, on the other hand, has not been specially

¹ The term *deliquescence* refers to the process of absorbing moisture from the air so that a salt becomes moist, or even dissolves in the moisture it has absorbed from the air. *E.g.* expose potassium carbonate to the atmosphere by placing some crystals in a small dish and note the result. The term *efflorescence* refers to the formation of a crust—generally white—on the surface of a body. The phenomenon is—very often—due to the loss of water from certain crystalline salts. *E.g.* expose crystals of washing-soda to a dry atmosphere.

purified and hence is sold at a cheaper rate than the "chemically pure" chemicals. Purification is an expensive operation, and the cheaper commercial reagents¹ are used whenever specially purified materials are not required. Some hold that "perfectly pure substances are unknown." This is probable, but to establish the proposition, we should be involved in a metaphysical discussion, and we might be led to say with A. Laurent: "Chemistry is the science of substances which do not exist."

The effect of traces of impurity on the properties of a compound.—It may be well to emphasize, just here, that sometimes a minute trace of impurity is of vital importance. Some reactions proceed quite differently in the presence, and in the absence of traces of impurity. The properties of many substances, too, are modified in a remarkable manner by small traces of impurity. H. Vivian says that $\frac{1}{1000}$ part of antimony will convert the best selected copper into "the worst conceivable"; Lord Kelvin says that the presence of $\frac{1}{1000}$ part of bismuth in copper would reduce its electrical conductivity so as to be fatal to the success of the submarine cable; and W. R. Roberts-Austen says that $\frac{1}{500}$ part of bismuth in gold would render gold useless, from the point of view of coinage, because the metal would crumble under pressure in the die. That such minute proportions of extraneous matter should be capable of communicating sensible properties of a definite character to the body with which they are mixed is, in the opinion of J. F. W. Herschel, "one of the most extraordinary facts that has appeared in chemistry."

§ 7. Dalton's Law of Multiple Proportions.

If Dalton's hypothesis of multiple proportions be found correct, we shall have to regard it as the greatest advance chemistry has yet made towards its development into a science.—J. J. BERZELIUS, 1811.

The formation of chemical compounds is not a capricious and fortuitous process, but it proceeds in an orderly fashion. Chemical combination is restricted to certain fixed proportions of matter. These limitations appear to have been prescribed by nature as part of her scheme in building the material universe. This fact arrested the attention of J. Rey in 1630. Rey's conclusion that in the calcination of the metals "nature has set limits which she does not overstep," agrees with many facts; but there are certain limitations. If one gram of lead be calcined for a long time at 500°, never more than 1.103 gram of a red powder—red lead—is obtained. Here, 64 grams of oxygen correspond with 621 grams of lead. If the lead be calcined at about 750°, one gram of lead will not take up more than 0.078 gram of oxygen to form a yellow powder—litharge; otherwise expressed, 64 grams of oxygen correspond with 828 grams of lead. Here then nature has set *two* limits; lead forms at least two definite oxides—a red oxide stable at a dull red heat, and a yellow oxide stable at a bright red heat. The relative proportions of lead and oxygen in the two oxides are as follows:

	Oxygen.	Lead.
Red oxide (red lead)	64	621 = 207 × 3
Yellow oxide (litharge)	64	828 = 207 × 4

¹ The terms *reagents* and *chemicals* are applied to the substances used in chemistry for producing special reactions with other substances. The term "reagent" is more particularly used in analytical work.

This means that for a given weight of oxygen, the yellow oxide has four-thirds as much lead as the red oxide. Similarly, carbon forms two well-defined oxides called respectively carbon monoxide and carbon dioxide. In these we have :

	Oxygen.	Carbon.
Carbon dioxide	8	$3 = 1 \times 3$
Carbon monoxide	8	$6 = 2 \times 3$

At least five oxides of nitrogen are known. In these, the relative proportions of nitrogen and oxygen are as follows :

	Nitrogen.	Oxygen.
Nitrogen monoxide	14	$8 = 1 \times 8$
Nitrogen dioxide	14	$16 = 2 \times 8$
Nitrogen trioxide	14	$24 = 3 \times 8$
Nitrogen tetroxide	14	$32 = 4 \times 8$
Nitrogen pentoxide	14	$40 = 5 \times 8$

These five compounds of the same elements united in different proportions form a series of substances so well marked and contra-distinguished that it is questionable if the most acute human intellect would ever have guessed that they contained the same constituents. Starting from the compound with the least oxygen, we see that for every 14 grams of nitrogen, the amount of oxygen increases by steps of 8 grams. Accordingly, in all five compounds of nitrogen and oxygen, the masses of nitrogen and oxygen are to one another as $m \times 14 : n \times 8$, where m and n are whole numbers. Hundreds of cases equally simple might be cited. Similar facts led J. Dalton (1802-4) to the generalization now called the law of multiple proportions : when one substance unites with another in more than one proportion, these different proportions bear a simple ratio to one another.

There is no difficulty in tracing the "simple ratio" $m : n$ in the cases which precede, but it is not always easy to detect the *simplicity* of this ratio in perhaps the larger number of cases. For instance, the ratio $m : n$ for compounds of carbon and hydrogen passes from 1:4 in methane, up to 60:122 in dimyricyl, and still more complex cases are not uncommon. Still, the law is considered to be so well founded that it can be applied to predict the composition of compounds which have never been prepared. Thus, if an oxide of nitrogen containing rather more oxygen than nitrogen pentoxide be made, we may predict that it will contain $6 \times 8 = 48$ parts of oxygen for every 14 parts of nitrogen by weight. Again, if a substance be found to contain oxygen and nitrogen, not in the proportion 14:8 or a multiple of 8, it is in all probability a mixture, not a true compound. Thus, air contains oxygen and nitrogen, but the proportions of nitrogen to oxygen is as 14:4.29. This is usually given along with other circumstantial evidence to show the probability that air is a mixture and not a chemical compound.

We might easily be led to reason in a vicious circle (*in circulo probando*) by a rigid application of the so-called multiple proportion law. A salt dissolves in water in all proportions up to a certain limiting value. The process of solution, in some cases, seems to be otherwise indistinguishable from chemical combination. It is sometimes said that the process of solution cannot be a case of chemical combination because there are no signs of abrupt *per saltum* changes characteristic of combination* in

multiple proportions. One writer has said: "Efforts have been made to find compounds which do not conform to the laws of chemical combination, but all attempts have resulted in failure;" another has said: "The law of multiple proportions has been tested by the analysis of thousands of compounds, and, like the law of constant proportions, it is one of the perfect laws from which no deviation has been discovered." It will be observed that if exceptions to the laws of chemical combination were discovered, chemists would refuse to call them compounds, and the quest for exceptions must therefore end in failure. Whether a solution is or is not a case of chemical action cannot be decided by an appeal to the laws of chemical combination; nor can the appeal to experiment establish or refute the laws of constant and multiple proportions. More bluntly expressed: a prejudice in favour of the generalization in question may warp the judgment to such an extent as to lead to a denial of the possibility of contradictory phenomena. Such a perversion of the judgment must be detrimental to the progress of science. Hence the danger of cherishing a blind faith in our so-called "laws of nature" which are often conventional definitions. (*Cf.* p. 570.)

§ 8. Richter's Law of Reciprocal Proportions.

After long and painful centuries of continuous effort, chemistry has discovered that the elements combine with one another in definite and unchanging ratios of quantity; and that, when their compounds are decomposed, they yield up those identical ratios.—S. BROWN (1843).

Between 1810 to 1812, J. J. Berzelius published the results of a careful study of the quantitative relations of some of the elements. He found that 100 parts of iron, 230 parts of copper, and 381 parts of lead are equivalent, for they unite with 29.6 parts of oxygen, forming oxides, and with 58.73 parts of sulphur, forming sulphides. Hence, since 58.73 parts of sulphur and 29.6 parts of oxygen unite respectively with 381 parts of lead, then, if sulphur and oxygen unite chemically, 58.73 parts of sulphur will unite with 29.6 parts of oxygen, or, taking the law of multiple proportions into consideration, with some simple multiple or submultiple of 29.6 parts of oxygen. In confirmation, Berzelius found that in sulphur dioxide, 58.73 parts of sulphur are united with 57.45 parts of oxygen. The difference between $2 \times 29.6 = 59.2$ and 57.45 is rather great, but some of the methods of analysis were crude in the time of Berzelius, and very much closer approximations—very nearly 1 in 50,000—have been obtained in recent years.

J. B. Richter, some twenty years before Berzelius' work, proved that a similar relation held good for the combination of acids and alkalis. Berzelius extended Richter's law¹ to combinations between the elements. The above relations are included in the generalization sometimes called the law of reciprocal proportions, or the law of equivalent weights. The weights—multiple or submultiple—of the various elements which react with a certain fixed weight of some other element, taken arbitrarily as a standard, also react with one another. If two substances, A and B, each combines with a third substance C, then A and B

¹ C. F. Wenzel, 1777, is sometimes said to be the father of this generalization. This, however, appears to be an historical error.

can combine with each other only in those proportions in which they combine with C, or in some simple multiple of those proportions. The laws of constant, multiple, and reciprocal proportions are wonderful examples of the beauty and harmony of nature; and yet, we have hints that these are but symbols of a sublimer generalization which, when discovered,

Will make one music as before
But vaster.

If a compound be formed by the union of two elements A and B, it is only necessary to find the proportion in which a third element C unites with one of the two elements, say A, to determine the proportions in which C unites with B. These numerical relations come out very clearly by comparing the proportions in which the different members of a series of elements, selected at random, combine with a constant weight of several other elements. Suppose the analysis of a substance shows that its ingredients are not in those proportions which we should expect from the known combinations of each of its components with another substance, we might safely infer that the substance analyzed is a mixture, and not a single compound.

§ 9. Combining, Reacting, or Equivalent Weights.

Since it is already settled for us by custom that quantities of different substances are to be called equal when or because they are equivalent gravimetrically, we have no choice but also, from the chemical point of view, to call those quantities of substance equal which interact in single chemical changes.—E. DRIVERS, 1902.

The following numbers represent the results obtained by the chemical analysis of a number of substances selected at random:

	Per cent.	Per cent.
Silicon dioxide	Silicon 46·93 ;	Oxygen 53·07
Hydrogen chloride	Hydrogen 2·76 ;	Chlorine 97·23
Magnesium chloride	Magnesium 25·53 ;	Chlorine 74·47
Water	Hydrogen 11·18 ;	Oxygen 88·81
Silver chloride	Silver 75·26 ;	Chlorine 24·74
Silver fluoride	Silver 70·05 ;	Fluorine 29·95

Analyses are generally calculated so that the sum of the constituents, all together, is 100 (per cent.) within the limits of experimental error. This is simply a convention of the analyst, for the results could be just as intelligibly summed to any other number. Taking any one of the elements as a standard, let us calculate what amount of each of the other elements will combine with a given quantity of the selected element. To save time, take oxygen = 8 as the standard. Starting with silicon, 53·07 parts of oxygen are combined with 46·93 parts of silicon. Consequently, we have the proportion

$$53\cdot07 : 8 = 46\cdot93 : x; \text{ or, } x = 7\cdot07$$

for silicon when oxygen = 8. Similarly, for water, hydrogen is 1·008 when oxygen is 8. Again, in hydrogen chloride, when hydrogen is 1·008, chlorine is 35·45; in silver chloride, silver is 107·88 when chlorine is 35·45;

when silver is 107·88, fluorine is 19; and when chlorine is 35·45, magnesium is 12·16. Collecting together the results of these calculations, we get

Oxygen	Silicon.	Hydrogen.	Chlorine.	Silver.	Fluorine.	Magnesium.
8	7·07	1·008	35·45	107·88	19	12·16

We have previously obtained a number of results for some metals for the standard $O = 8$ by a different process, and the number for magnesium obtained by an indirect process: Oxygen \rightarrow hydrogen (water) \rightarrow chlorine (hydrogen chloride) \rightarrow magnesium (magnesium chloride) gives the same result within the limits of experimental error as was obtained by a totally different process. Similar results are obtained in all cases, subject, of course, to the greater risk of experimental error when a long chain of compounds is involved. As a rule, there is no need to follow such an extended series as we have done here for fluorine and magnesium. Most of the elements unite directly with oxygen; and with the other elements, one intermediate step usually suffices.

We are therefore able to deduce the important generalization. The combining weight of the elements are specific constants, *i.e.* they change from element to element, but for each element the combining weight is fixed and invariable. Otherwise expressed: a number can be assigned to each element; this number—called the combining, reacting, or equivalent weight—represents the number of parts by weight of the given element which can enter into combination with 8 parts by weight of oxygen, or one part by weight of hydrogen. All combining weights are relative numbers, and they are conventionally referred to oxygen = 8, or hydrogen = 1. When an element unites with another element in more than one proportion, the higher proportions will always be simple multiples of the combining weights—one for each element. This is the so-called law of combining or reacting weights: when substances enter into chemical combination they always do so in quantities proportional to their combining weights.

If the combining weights of the elements are fixed, as they undoubtedly are, and since the elements can combine to form compounds which, in turn, can form compounds with other elements and with one another, it follows that the compounds themselves also have combining weights if they also can enter into chemical combination. Hence the so-called law of compound proportion—the combining weight of a compound body is the sum of the combining weights of its components. This deduction from the law of combining weights is as firmly established experimentally as the law of combining weights itself. The neutralization of acids by bases, and numerous chemical reactions, can be cited in illustration.

The experimental results stated on p. 21, Table I., raise the suspicion that *there is a difference between chemical and gravitational equality*. In the latter, equal quantities of the different forms of matter are represented by equal weights; whereas, in a chemical sense, equal quantities of matter are the weights or masses of different forms of matter which unite with one another chemically. Consequently, chemical union may be regarded as a measure of the amounts of the different forms of matter which are chemically equivalent. Chemical equality is thus as clearly defined as gravimetric equality. The former is a measure of chemical and the latter a measure of physical phenomena; the latter is wholly independent

of, and the former mainly dependent upon, the nature of the substances compared.

§ 10. The Law of the Perdurability of Matter.

The annihilation of matter is unthinkable for the same reason that the creation of matter is unthinkable, the reason namely that nothing cannot be an object of thought.—H. SPENCER.

It will be remembered that Lavoisier (1774) heated tin with air in a closed vessel and found that the weight of the whole system, before and after the calcination of the tin, was the same, thus showing that the whole system had neither gained nor lost in weight. This experiment is mentioned because it emphasizes, very well, the fact that in spite of the most painstaking care, every time all the substances taking part in a chemical reaction are weighed before and after the change, there is no sign of any alteration in the quantity of matter. This fact is sometimes called the law of the indestructibility of matter. As Democritus has said, nothing can never become something, nor can something become nothing—*ex nihilo nihil fit, et in nihilum nihil potest reverti*. The principle of the indestructibility of matter was tacitly assumed by many old investigators. A. L. Lavoisier is generally supposed to have first demonstrated the law in 1774 by experiments like that cited above, but the law was definitely enunciated in 1756 by M. W. Lomonossoff, and the law must have been at the back of J. Black's mind when he worked on the alkaline earths (*q.v.*) in 1755.

The chemist's law of "the indestructibility of matter" really means that the total *weight* of the elements in any reacting system remains constant through all the physical and chemical changes it is made to undergo. The observed facts are better generalized as the law of persistence of weight; no change in the total weight of all the substances taking part in any chemical process has ever been observed. If A and B represent respectively the weights of two elements which take part in a chemical reaction, producing the weights M and N of two other substances, the law of persistence of weights states that $A + B = M + N$ where "+" means "together with," and "=", produces. If the weight of one of these four substances be unknown, it can be computed by solving the equation. Chemists constantly use this principle in their work; for, as Lavoisier said: "Experiments can be rectified by calculations, and calculations by experiments. I have often taken advantage of this method in order to correct the first results of my experiments, and to direct me in repeating them with proper precautions."

When faith in magic was more prevalent than it is to-day, many believed that by some potent incantation or charm, matter could be called out of nothingness, or could be made non-existent.¹ Superficial observation might lead to the belief that a growing tree, the evaporation of water, and the burning of a candle prove the creation and the destruction of matter, but a careful study of these and innumerable other phenomena,

¹ H. Spencer considers that all the so-called experimental proofs by weighing tacitly assume the object being proved, since weighing implies that the matter forming the weights remains relatively unchanged in quantity; or, as H. S. Redgrove pointed out, weight measures matter because matter is indestructible, and matter is indestructible because weight measures matter.

has shown that the apparent destruction of matter is an illusion. Matter may change its state as when liquid water is vaporized, and when a candle is burnt. In the case of a growing tree, the nutrition the tree receives from the soil and from the air (carbon dioxide) is overlooked. Fig. 7 illustrates an instructive experiment which is commonly used to show that the apparent destruction of matter in the burning of a candle is illusory. A candle, *A*, is fixed on one pan of a balance below a cylinder *B*. A piece of coarse wire gauze, *C*, is fixed in the lower part of the cylinder *B*. The

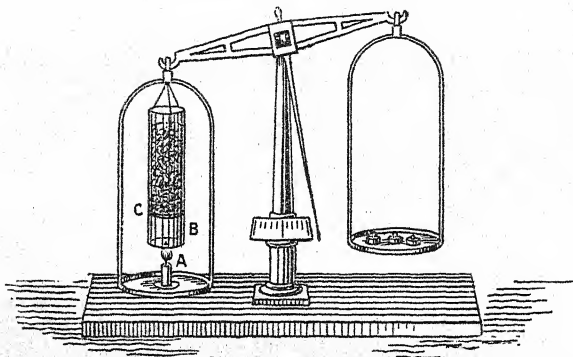


FIG. 7.—Apparent Increase in Weight during Combustion
(after H. E. Roscoe and C. Schorlemmer).

wire gauze supports a few lumps of quicklime¹ on which rests a mixture of granulated soda lime and glass wool—the latter to prevent the soda lime clogging the tube. Weights are added to the right scale pan until the beam of the balance is horizontal. The candle is lighted. The gases rising from the flame pass through the cylinder *B*—and the products of combustion are absorbed by the soda lime. In 3 or 4 minutes the pan carrying the candle is depressed as illustrated in the diagram. The increase in weight is due to the fixation of the products of combustion by the soda lime. The products of combustion are formed by the combination of the carbon and hydrogen of the candle with the oxygen of the air. The oxygen of the air was not weighed in the first weighing.

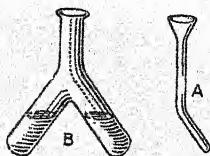


FIG. 8.—Landolt's Tube.

Every time a chemical reaction takes place in a closed vessel, which permits neither the egress nor the ingress of matter, the total weight remains unchanged within the limits of experimental error. The more carefully the experiments are made, the more nearly do the values approach identity. Both A. Heydweiller (1901) and H. Landolt (1893) have tried to find if a loss in weight occurs during chemical action. Their experiments may be illustrated by introducing a solution of silver nitrate into one limb of the Λ -shaped tube *B*, Fig. 8, by means of the funnel *A*, and a solution of potassium chromate in the other limb. The tube is then weighed and tilted so as to mix the solutions and start the reaction. There

¹ To prevent water dripping on to the flame.

is no difference in the weight of the tube, before and after the reaction, within the limits of the experimental error: ± 0.000030 gm. In Landolt's experiments the opening of the tube *B* was sealed up before the solutions were mixed. Other pairs of solutions are: a solution of potassium iodate slightly acidulated with hydrochloric acid and potassium iodide; lead acetate and sodium sulphide; acidulated potassium chromate and sodium sulphite, etc. After an examination of fifteen different reactions, Landolt (1909) failed to detect a variation in weight, and "since there seems no prospect of pushing the precision of the experiments further than the degree of exactness attained, the experimental proof of the law may be regarded as established." The law of the persistence of weight can thus be stated: A variation in the total weight of the substances taking part in chemical reactions, greater than the limits of experimental error, has never been detected.

It is quite conceivable that the weight of the iron in, say, magnetic oxide of iron might *appear* to be greater than the same amount of iron in, say, potassium ferrocyanide, because of the effect of the earth's magnetic field upon the former. But if such an effect were observed, it would not interfere with our faith in the law as soon as the disturbing effect was recognized.

§ 11. The Atomic Theory.

It seems probable to me that God in the beginning formed matter in solid, massy, hard, impenetrable, movable particles, of such sizes and figures, and with such other properties, and in such proportion to space, as most conduced to the end for which He formed them; and that these primitive particles, being solids, are incomparably harder than any porous body compounded of them, even so very hard as never to wear or break in pieces; no ordinary power being able to divide what God Himself made one in the first creation. . . . The changes of corporeal things are to be placed only in the various separations and new associations and motions of these permanent particles. . . . These principles I consider not as occult qualities, but as general laws of nature by which the things themselves are formed; their truth appearing to us by phenomena, though their causes be not yet discovered.—ISAAC NEWTON.

The four laws of chemical combination: (1) the persistence of weight; (2) the law of constant composition; (3) the law of multiple proportions; and (4) the law of reciprocal proportions, summarize observed facts. They exist quite independently of any hypothesis we might devise about their inner meaning; but we have an intuitive feeling that there must be some peculiarity in the constitution of matter which will account for the facts.

An atom is the unit of chemical exchange.—Chemists in imagination have invested matter with a granular structure. Matter is supposed to be discrete, and built up of corporeal atoms. The imagination can subdivide matter indefinitely; the chemist says that however true this may be, nothing less than an atom ever takes part in a chemical reaction. The atom is the limiting size so far as chemical combination is concerned. An atom cannot be subdivided by any known chemical process. What A. Kekulé wrote in 1867 applies equally well to-day, in spite of some interesting though abortive attempts to eliminate atoms from chemistry. Should the progress of chemistry lead to a different view of the constitution of matter, it will make little alteration to the chemist's atom. The chemical

atom will always remain the chemist's unit. "As a chemist," wrote Kekulé, "I believe that the assumption of atoms is not only advisable but absolutely necessary provided that the term be understood to denote those particles of matter which undergo no further division in chemical transformations."

Compare this hypothesis with observation. Fix the attention on the facts: Elements combine with one another either in amounts which correspond with their combining weights (law of constant composition), or with multiples of their combining weights (law of multiple proportions). Otherwise expressed, definite amounts of matter—the atoms—corresponding with the combining weights act as chemical units. Reactions between different elements are reactions between these units. Atoms of the same element all have the same constant weight,¹ and atoms of different elements have different combining weights. All this is in agreement with the law of constant combining weights.

Fractions of an atom do not take part in chemical changes.—The proportions in which one element combines with another can alter only by steps one atom at a time. 1, 2, 3, . . . atoms of one element can combine with 1, 2, 3, . . . atoms of another element. This is but one way of stating the laws of multiple and reciprocal proportions. The weight of an atom of each element is a constant quantity, and therefore elements can only combine with each other in certain constant proportions or in multiples thereof. The atoms of the elements are the units from which nature has fashioned all the different varieties

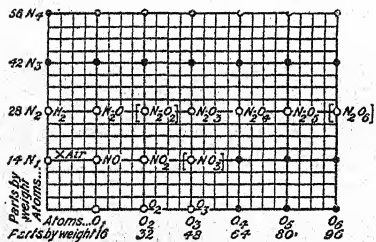


FIG. 9.—Law of Multiple Proportions.

of matter in the universe. One atom of mercury unites with one atom of oxygen to form mercuric oxide. If two atoms of mercury united with one atom of oxygen, the result would not be mercuric oxide, but some other oxide of mercury—if otherwise, the law of constant composition would be false. As a matter of fact, such a compound is known, but it is mercurous oxide. Mercurous oxide has its own specific properties which are different from those of mercuric oxide. We thus adopt the view of J. B. Dumas and of M. Faraday that "whether matter be atomic or not, this much is certain, granting it be atomic, it would behave in chemical transformations as it does now."

The law of multiple proportions.—The idea embodied in this law can be neatly illustrated by means of squared paper. Let the abscissæ, Fig. 9, represent weights of oxygen, and the ordinates weights of nitrogen which enter into combination with the oxygen. If, for convenience, we take oxygen = 16, and nitrogen = 14 as standard weights for the respective atoms, the known unary and binary compounds of these elements appear at the points of intersection of lines representing 14, 28, 42 . . . parts

¹ It might here be added that W. Crookes (1887) found it expedient to assume as a working speculation that the atoms of the element yttrium "differ probably in weight, and certainly in the internal motions they undergo." There is, however, a proof of this in the existence of the so-called isotopes discussed later on.

by weight of nitrogen, and 16, 32, 48 . . . parts by weight of oxygen. Let the symbol N represent one atom of nitrogen; N_2 , two atoms of nitrogen, etc.; and let O represent one atom of oxygen, O_2 , two atoms of oxygen, etc. Then the symbol N_2O , NO, etc., may be used to represent compounds containing the indicated number of atoms of each element. The known compounds of these two elements are represented by symbols "o" in the diagram. If the law of multiple proportions perfectly describes nature's *modus operandi*, and no other disturbing influence be at work, we can represent possible,¹ but yet undiscovered, compounds of nitrogen and oxygen by points "•"; and compounds representable by the intermediate positions in the diagram are impossible. Hence air, whose composition is represented by the cross "x" in the diagram, is not a chemical compound.

Atomic weights are relative.—We can express the combining weights of the atoms in terms of any unit we please; it is quite immaterial whether a gram or a ton be imagined. In dealing with combining or atomic weights, the conception of absolute quantity is quite irrelevant. Given sufficient oxygen, 100 tons, kilograms, pounds, grams, or grains of mercury will give 108 tons, kilograms, pounds, grams, or grains, respectively, of mercuric oxide—no more, no less. We do not know the actual weights, but we do know the relative weights. The indirect estimates which have been made show that the weight of an atom of hydrogen is not far from 12.8×10^{-20} milligram. Accepting this estimate, the Table of Atomic Weights enables the weight of an atom of any other element to be calculated. Thus, the chlorine atom weighs $35.5 \times 10^{-20} \times 12.8$ mgrm.

The history of the atomic hypothesis.—The ancient philosophers of the East—India, Greece, Italy, etc.—made many quaint guesses at the constitution of matter. Among these guesses, we find one taught by Kanáda (the founder of a system of Hindu philosophy) long prior to the rise of Grecian philosophy. The same guess was made by Democritus, Leucippus, and Lucretius, and their guess lives, more or less modified, in modern chemistry. These philosophers seem to have taught: (1) matter is discrete; (2) all substances are formed of atoms which are separated from one another by void space; (3) the atoms are in constant motion; and (4) motion is an inherent property of the atoms. The atoms were supposed to be too small to be perceived by the senses, and they were further supposed to be eternal, indestructible, and unchangeable. Atoms differed from each other in shape, size, and mode of arrangement, and the properties of all substances were supposed to depend upon the nature of the constituent atoms and the way the atoms were arranged. So far as the experimental evidence available to the Grecian philosophers in support of this particular guess is concerned, its long life—in the form of the chemist's atomic theory—can only be attributed to chance. The modern theory, unlike the old speculation, is based upon the observed laws of chemical change, and it cannot stand apart from them.

Many thinkers—Francis Bacon, René Descartes, Pierre Gassendi, Robert Boyle, Robert Hooke, John Mayow, etc.—were more or less partial

¹ The symbols in brackets in the diagram represent compounds which we are not so sure about. It is also conceivable that at some future time it may be necessary to extend the diagram upwards and to the right. A similar diagram for the compounds of carbon and hydrogen indicates scores of known compounds.

to a theory of atoms. Isaac Newton (1675) tried to explain Boyle's law on the assumption that gases were made up of mutually repulsive particles; and the above quotation shows that he also referred chemical changes to different associations of the atoms. M. W. Lomonosoff, also, had a fairly clear concept of the atomic structure of matter in 1748; while Bryan Higgins (1776) and William Higgins (1789) explained the constant composition of salts, with more or less confidence, in terms of the atoms. Bryan Higgins appears to have held the view that two different atoms combine in the proportions of 1:1, and in that proportion only; while William Higgins imagined a combination in multiple proportions, but believed that the combination 1:1 was the most stable.

Dalton's atomic hypothesis.—It is thus impossible to say who invented the atomic theory, because it has grown up with chemistry itself. In the work of William Higgins the hypothesis was little more than an inanimate doctrine. It remained for Dalton to quicken the dead dogma into a living hypothesis. John Dalton (1801) employed the atomic hypothesis to explain the diffusion of gases, and later (1803) based an hypothesis of the structure of matter and of chemical combination upon the following postulates, which may be regarded as a very brief statement of the so-called Dalton's atomic theory:

1. Atoms are real discrete particles of matter which cannot be subdivided by any known chemical process.
2. Atoms of the same element are similar to one another, and equal in weight.
3. Atoms of different elements have different properties—weight, affinity, etc.
4. Compounds are formed by the union of atoms of different elements in simple numerical proportions—1:1; 1:2; 2:1; 2:3; etc.
5. The combining weights of the elements represent the combining weights of the atoms.

The hypothesis of Dalton respecting atoms, and more particularly atomic weights, is not quite that which prevails in modern chemistry. John Dalton considered the atom to be indivisible, and this is expressed in his aphorism: "Thou knowest no man can split an atom." T. Graham defined the atom not as a thing which cannot be divided but as one which had not been divided. The modern idea is that while the atom is perdurable in chemical changes it may be and probably has been resolved into component parts. A formidable meta-chemistry has been elaborated, and is wholly based on the assumption that an atom is a complex system of sub-atoms or electrons.

The defect in Dalton's atomic theory.—According to the atomic theory: an atom is the smallest particle of an element which can enter into or be expelled from chemical combination.¹ How is the "smallest combining weight" of an atom to be fixed? In carbon monoxide, for example, we have oxygen and carbon in the following proportions by weight:

Oxygen : Carbon = 8 : 6

and in carbon dioxide

Oxygen : Carbon = 8 : 3

¹ If we think of the derivation of the word *atom*—from the Greek *ατομος* (*temno*), I cut—"that which cannot be subdivided," we must add "chemically." But our definition of the atom says nothing about subdivision; nor about the ultimate nature of the atom. The term "atom" was once used to represent the "smallest interval of time," a "moment."

What is the atomic weight of carbon if the atomic weight of oxygen is 8? Obviously, the evidence now before us would be consistent with many different views. Carbon monoxide may be a compound of one oxygen atom with two carbon atoms each with a combining weight of 3; or a compound of one oxygen atom with one carbon atom with a combining weight of 6. In the latter case, carbon dioxide is a compound of one carbon atom combining weight 6 with two oxygen atoms, and the same combining weights would have been obtained if any number n of carbon atoms were combined with $2n$ oxygen atoms. Similar difficulties arise when we apply the idea of atoms so far developed to other combinations of the elements. There is, therefore, some confusion. The concept of the atom becomes more or less indistinct and vague when the attempt is made to develop a consistent system on the basis of the atomic hypothesis as propounded by Dalton. Dalton's theory is defective because it lacks a standard for fixing the atomic weights of the different elements.

It may be perfectly true, as Lord Kelvin has pointed out, that "the assumption of atoms can explain no property of a body which has not previously been attributed to the atoms," but the assumption has none the less proved an invaluable aid in forming mental concepts of the different phases of a chemical reaction; and it has enabled chemists to successfully anticipate the results of experimental research. A. R. A. Smith said in 1884: "We believe in atoms because, so far as we can see, nature uses them." The greater the number of facts consistently explained by one and the same theory the greater the probability of its being true. The overwhelming mass of circumstantial evidence, direct and indirect, which modern chemistry and physics offers, has justified the faith of Dalton; and demonstrated the existence of veritable atoms. Still further, in the concluding chapters of this work, evidence will be indicated which shows that the atoms have a complex structure.

§ 12. The Language of Chemistry.

However certain the facts of any science, however just the ideas derived from these facts, we can only communicate false or imperfect impressions to others, if we want words by which these may be properly expressed.—A. L. LAVOISIER.

The nomenclature of a science, that is, the group of technical terms peculiar to that science, is of vital importance. It is virtually impossible to separate the nomenclature from the science itself. Lavoisier emphasized the importance of this in his classical *Traité Élémentaire de Chimie* (1789). Every science consists of three things: (1) the facts which form the subject-matter; (2) the ideas represented by those facts; and (3) the words in which those ideas are expressed. "Like three impressions of the same seal, the word ought to produce the idea; and the idea ought to be a picture of the fact."

Special technical words are employed to fix and describe the ideas and principles of chemistry—as of all other sciences; and words are borrowed from colloquial everyday language, and used with a specific meaning. In any case technical terms should be precise and clear, and not tainted with ambiguity and vagueness. Such technical terms form part of the current language of chemistry. Words with a variety of meanings are a

prolific source of error and confusion, and they have thus provoked numberless controversies. Specially invented terms are less liable to misapprehension than terms adapted from current words. However strange the terms may appear at first, they soon grow familiar to the ear, and they can then be used without effort. W. Whewell has pointed out, very aptly, that "technical terms carry the results of deep and laborious research. They convey the mental treasures of one period to the generations that follow, and laden with this, their precious freight, they sail safely across the gulfs of time in which empires have suffered shipwreck, and the language of common life has sunk into oblivion"—witness: some of the terms used in the chemistry of to-day were coined by the early Arabian chemists—alcohol, alkali, borax, elixir, etc.

Naming the elements.—A great number of the elements have been christened with names derived from Greek roots. *E.g.*, *iodine*—from its violet vapour; *chlorine*—from its green colour; *chromium*—from the colour of its compounds; *rhodium*—from the rose colour of its salts; *osmium*—from its smell; *helium*—from its occurrence in the sun; *argon*—from its indifference to chemical reagents, etc. Other elements have been named more or less capriciously; thus some elements are named after particular localities—*strontium*, from Strontian (in Scotland); *ruthenium*, from Ruthenia (Russia); *yttrium*, *ytterbium*, *erbium*, and *terbium* are all derived from Ytterby (in Sweden); *palladium* is a name given in honour of the discovery of the planetoid Pallas; *uranium* in honour of the discovery of the planet Uranus; *beryllium* is derived from the name of the mineral beryl; *zirconium*, from the mineral zircon; *platinum*, from the Spanish "plata," silver; *victorium*, from Queen Victoria; *thorium*, from "Thor," the son of Odin, a god in Scandinavian mythology; *vanadium*, from a Scandinavian goddess, Vanadis; *tantalum*, from Tantalus in Grecian mythology; and *niobium*, from Niobe, daughter of Tantalus.

Symbols.—The old alchemists used to represent different substances by symbols. For example, gold was represented by the symbol ☉ or ✱, for the sun; silver, by ☾, the moon; etc. Lavoisier used the symbol ∇ for water; Ⓐ for oxygen; etc. Dalton made a step in advance by representing the atoms of the elements by symbols, and combining these symbols so as to show the elements present in a compound. Thus, ☉ represented hydrogen; ○ oxygen; ● carbon; etc. Water was represented by ☉○; carbon monoxide by ○●; carbon dioxide by ○●○; etc. These symbols have all been abandoned. They are too cumbrous. To-day we follow J. J. Berzelius' method, suggested in 1811, and use one or two letters from the recognized name of the element to represent, any particular element.¹ Thus, O represents oxygen; H, hydrogen; C, carbon; N, nitrogen; Cl, chlorine; etc. The names of ten elements start with C, and to prevent the possibility of confusion, a second leading letter is selected either from the name, or from the alternative Latin name of the element. Thus, C (carbon), Cb (columbium), Ca (calcium), Cd (cadmium), Ce (cerium), Cl (chlorine), Co (cobalt), Cr (chromium),

¹ Unfortunately some elements have not yet been christened with a name recognized by all. Niobium—symbol Nb—and columbium—symbol Cb—are two different names for one element; glucinum—symbol Gl—and beryllium—symbol Be—are two different names for another element.

Cs (caesium), and Cu (cuprum, copper). The elements with alternative Latin names are symbolized: Sb for antimony (Lat. stibium); Cu for copper (Lat. cuprum); Au for gold (Lat. aurum); Fe for iron (Lat. ferrum); Pb for lead (Lat. plumbum); Hg for mercury (Lat. hydrargyrum); K for potassium (Lat. kalium); Na for sodium (Lat. natrium); and Sn for tin (Lat. stannum).

The evolution of chemical nomenclature.—Up to near the close of the 18th century, no systematic attempt had been made to name chemical substances in such a way as to indicate their composition. The names then in vogue were more or less arbitrary, for they were relics of alchemical terms—e.g. *crocus martis*—or derived from their discoverer—e.g. *Glauber's salt*—or based on some superficial resemblance furnishing what J. B. Dumas called the language of the kitchen. Thus, antimonious chloride was called *butter of antimony* because of its buttery appearance; zinc chloride, *butter of zinc*; and arsenic chloride, *butter of arsenic*. These three substances were classed together with butter from milk. Similarly with oil of vitriol, olive oil, etc.; spirits of wine, spirits of salt, etc. T. Bergman and G. de Morveau simultaneously and independently attempted to devise a more complete system of naming chemical compounds. A. L. Lavoisier presented a report to the French Academy and terms like “ic” and “ate,” “ous” and “ite” were employed. J. J. Berzelius followed up the subject, and inaugurated the system which is virtually that employed to-day.

Naming the compounds.—Each element forms with other elements a group of compounds which are said to *contain* the respective elements, because the elements in question can be obtained unchanged from the compounds. Consequently every compound has an elementary or ultimate composition. Compounds are symbolized by joining together the letters corresponding with the different elements in the compound. Thus, HgO represents mercury oxide, a compound of mercury and oxygen. When only two elements are united to form a compound, the name of the second element is modified so that it ends in *ide*.

The symbol for the element also represents one of its atoms. If more than one atom is present in a compound, a small figure is appended to the bottom ¹ right-hand corner of the symbol for an atom of the element, to indicate the number of atoms present. Thus “H₂O” represents a molecule of water, i.e. a compound containing two atoms of hydrogen and one of oxygen; “CO” represents a molecule of carbon monoxide—a compound containing one atom of carbon and one atom of oxygen; “Na₂CO₃” represents a molecule of sodium carbonate—a compound containing two atoms of sodium, one atom of carbon, and three atoms of oxygen. A letter affixed in front of a group of symbols represents the number of times that group occurs in the given compound. Thus crystallized sodium carbonate is symbolized: Na₂CO₃.10H₂O. This means that this compound contains the equivalent of one Na₂CO₃, and ten equivalents of the group H₂O.

Compounds of one element with oxygen are called oxides, and the process of combination is called oxidation. When an element forms more than one oxide, a Greek numerical suffix is often prefixed to the word “oxide.” Thus, SO₂ is sulphur dioxide; SO₃, sulphur trioxide;

¹ In France, generally at the top.

CO, carbon monoxide; CO₂, carbon dioxide; PbO, lead monoxide; PbO₂, lead dioxide or lead peroxide.

Some of the commoner prefixes are :

	1	2	3	4	5	6	
Latin . .	Uni-	Bi-	Ter-	Quadri-	Quinque-	Sexa-	
Greek . .	Mono-	Di-	Tri-	Tetra-	Penta-	Hexa-	
	7	8	9	10	11	12	
Latin . .	Septa-	Octo-	Novem-	Decem-	Undecem-	Duodecem-	
Greek . .	Hepta-	Octo-	Ennea-	Deka-	Endeka-	Dodeka-	
	Half	One-third	One-fourth	One-fifth	Whole	Equal	Many
Latin . .	Semi-	trita	tetrata	pentita	Omni-	Equi-	Multi-
Greek . .	Hemi-				Holo-	Homo-	Poly-

It is considered bad style to mix Latin and Greek root words and prefixes. Consequently we usually try to keep Greek with Greek, and Latin with Latin. Thus, we say "diatomic," not "biatomic"; "bimolecular," not "dimolecular"; "bivalent," not "divalent"; and "bivariant," not "divariant"; because "atomic" is derived from a Greek word, while "molecular," "variant," and "valent," are derived from Latin words. There are, however, many hybrids universally recognized; e.g. millimetre, centimetre, etc. Monovalent, divalent, etc., are also used at times in spite of their hybrid character. We cannot, therefore, always be "purists" without defying custom, which, as Horace has said, decides the language we must use.

Sometimes the termination *-ic* is affixed to the name of the metal for that oxide which contains the greater proportion of oxygen, and *-ous* for the oxide containing the lesser proportion of oxygen.¹ For instance, SnO is either stannous oxide, or tin monoxide; FeO is ferrous oxide; and Fe₂O₃ ferric oxide. The last-named method of naming the compounds is not always satisfactory when the elements form more than two compounds. To get over the difficulty, a prefix *hypo-* (meaning "under," or "lesser") is sometimes added to the compound containing the least, and *per-* ("beyond," "above") is added to the one with the most oxygen. Thus,

Persulphuric acid	H ₂ S ₂ O ₈	Perchloric acid	HClO ₄
Sulphuric acid	H ₂ SO ₄	Chloric acid	HClO ₃
Sulphurous acid	H ₂ SO ₃	Chlorous acid	HClO ₂
Hyposulphurous acid	H ₂ S ₂ O ₄	Hypochlorous acid	HClO

The five nitrogen oxides—nitrogen monoxide, dioxide, trioxide, tetroxide, and pentoxide—would be awkwardly named by this system. Oxides like alumina—Al₂O₃; ferric oxide—Fe₂O₃, etc., are sometimes called *sesquioxides* (Latin, *sesqui*, one-half more).²

The nomenclature of inorganic chemistry is thus based upon the

¹ For historical reasons, the names of some compounds do not conform to this system because the affix "ic" was assigned to the compound first discovered, and the compounds subsequently discovered were named accordingly.

² The oxides can be roughly divided into two classes. Some oxides, with water, form acids, and others act as bases. It is not very easy, at this stage of our work, to draw a sharp line of demarcation between the two. The *acidic oxides* have a sour taste, and turn a solution of blue litmus red; the *basic oxides* turn a solution of red litmus blue, and have a soapy feel.

principle that the different compounds of an element with other elements can be named by a simple change in the beginning or termination of the word—witness ferric and ferrous oxides; and also by the addition of a numerical suffix showing the relative number of atoms of the corresponding element in its compounds. These little artifices, apparently trivial, are really important advances in the language of chemistry. The method has some defects, but when the necessity for a modification becomes acute, it will probably not be difficult to change. Language generally lags in the wake of progress.

Questions.

In the questions appended to this and subsequent chapters, I have omitted direct "quiz" questions such as, "What is the symbol of hydrogen?" "How is hydrogen prepared?" etc. Questions of this kind have a certain value in revision work, and good sets will be found in the brochure published as an "Appendix for Newth's Inorganic Chemistry"; in H. P. Talbot's "Study Questions—Inorganic Chemistry and Qualitative Chemical Analysis"; and in F. Jones' "Questions on Chemistry." Most of the questions here reproduced involve a little thought or work of comparison, and they will accordingly serve to emphasize special features which different examiners have thought to be important. Socrates long ago demonstrated the value of questions as an auxiliary in clarifying hazy notions.

1. Matter is said to be indestructible; what does this mean? What evidence is there that this is so?—*Science and Art Dept.*

2. When substances are brought into contact, how would you know whether they acted chemically on one another, or simply remained mechanically mixed? Describe minutely in illustration the combination of any two elements.—*Oxford Junior Locals.*

3. Distinguish as clearly as you can between changes in matter which are classed as chemical and those classed as physical. Which of the following do you think are chemical, and which are physical: (a) freezing ice cream; (b) souring milk; (c) burning a candle; (d) distilling water; (e) magnetizing iron; (f) electrolysis of a solution of copper sulphate?

4. Give a brief outline of the atomic theory, together with its history.—*Princeton Univ., U.S.A.*

5. Show how the facts summarized in the law of multiple proportions are explained by the atomic theory. Two oxides of a metal M contain respectively 22.53 and 30.38 per cent. of oxygen. If the formula of the first oxide be MO what will be that of the second?—*Sheffield Univ.*

6. In what relations do the elements stand to other single homogeneous substances, and such substances to other natural or artificial objects? Refer for illustration to quicksilver, water, gunpowder, a piece of granite, an orange, a picture.—*New Zealand Univ.*

7. State the law of multiple proportions. How would you proceed to prove the law experimentally?—*Aberystwyth Univ.*

8. John Dalton expressed and helped to establish the law of multiple proportions, and he also proposed an atomic theory. Explain why one of these contributions to chemical science was classed as a law while the other was not.—*American Coll.*

9. Define the Laws of Definite and Multiple Proportions. Show how the following analyses of three oxides of nitrogen illustrate the Law of Multiple Proportions:—

	A.	B.	C.
Nitrogen	63.65	46.68	25.94
Oxygen	36.35	53.32	74.06

Victoria Univ., Manchester.

10. Define a chemical change. Give instances of chemical changes occurring in nature. Describe experiments which show that matter is not lost in chemical changes.—*Owens Coll.*

11. Amplify the quotation : " Before everything a man of science should aim at being definite, clear, and accurate."

12. Describe *in detail* how you would separate the constituents of ordinary black gunpowder, and ascertain the percentage in the mixture of each constituent.
—*St. Andrews Univ.*

13. Criticize the experiment illustrated by Fig. 7 and point out where the result is inconclusive, and what would you suggest to overcome the objection ?
Hint : Note that soda-lime and quicklime increase in weight by mere exposure to atmospheric air.

CHAPTER III

WATER AND HYDROGEN

§ 1. Hydrogen—Preparation and Properties

THE attention of the early workers on chemistry was mainly directed to visible and tangible liquids and solids, while the "spirits," "fumes," "vapours," "airs," or gases, which escaped when certain substances reacted, were considered to be unwholesome effluvia, and were best avoided. To-day, a great deal of the student's time is occupied with the imperceptible and intangible gases ignored by the early workers. The study of gases began to occupy serious attention towards the end of the 18th century, so that, in 1779, although only eight were certainly known with respect to their composition, T. Bergman could complacently write: "During the last ten years, chemistry has not only soared into regions of invisible aerial substances, but has dared to explore the nature of these substances and to search into their constituent principles."

History.—It has been known for a very long time that an air or gas is produced when iron is dissolved in dilute sulphuric acid. T. B. Paracelsus, in the 16th century, described the action somewhat quaintly. He said that when the acid acts on iron "an air arises which bursts forth like the wind." J. B. van Helmont (c. 1609) described this gas as a peculiar variety of air which was combustible and a non-supporter of combustion, but his ideas were somewhat hazy, for he confused hydrogen with other gases, like methane and carbon dioxide, which do not support combustion. Priestley, and writers generally up to about 1783, used "*inflammable air*" as a general term to include this gas, as well as the hydrocarbons, hydrogen sulphide, carbon monoxide, and other combustible gases. H. Cavendish (1766) showed that the inflammable air produced by the action of dilute sulphuric or hydrochloric acid on metals like iron, zinc, and tin was a distinct definite substance, and A. L. Lavoisier (1783) called the gas "hydrogen."

The gas obtained from metallic iron is not very pure, and it possesses a distinct smell owing to the presence of hydrocarbon gases, etc., formed by the action of the acid on the carbon compounds associated, as impurities, with commercial iron. The solution remaining after the action of sulphuric acid on the iron when put aside in a cool place soon forms beautiful pale green crystals of ferrous sulphate. Their mode of formation, etc., is quite analogous with the process used in the preparation of zinc sulphate crystals, Fig. 5. Magnesium and aluminium furnish a fairly pure gas; with aluminium the acid should be warmed to start the reaction.

In these cases not only is hydrogen gas obtained, but crystals of magnesium sulphate and of aluminium sulphate can be obtained from the liquids in which the metals have been dissolved. The action of the acid on tin is rather slow; granulated zinc is used for general laboratory work, in the following manner:

Preparation in the laboratory.—Granulated zinc is placed in a two-necked Woulfe's bottle,¹ *A*, Fig. 10. One neck is closed air-tight by a one-hole rubber stopper fitted with a funnel tube, *B*, extending nearly to the bottom of the bottle; the other neck is fitted with a glass tube, *C*—delivery tube—bent as shown in the diagram. The delivery tube dips under the beehive, *D*, placed in a basin of water, *E*. The vessels *D* and *E* form a collecting, gas, or pneumatic trough.²

Pour some water through the funnel tube until the zinc is well covered; make sure that all the joints are air-tight, and that no escape of gas is possible other than through the delivery tube. Pour concentrated sulphuric acid, a little at a time, through the funnel tube until the gas begins to come off vigorously. The mixture of air and hydrogen gas first issuing

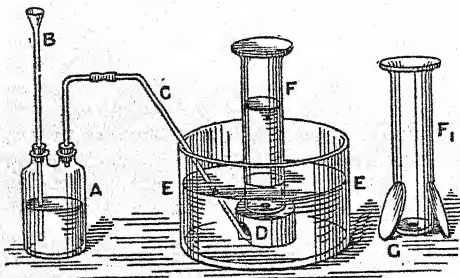


FIG. 10.—The Preparation of Hydrogen.

from the delivery tube is very explosive. It is therefore necessary to make sure that all the air has been expelled before the hydrogen is collected in the gas cylinders, or gas jars, *F*. Hence, invert a test tube full of water over the hole in the upper floor of the "beehive." When the tube is full of gas, apply a lighted taper to the mouth of the test tube. If the gas detonates, repeat the trial until the gas burns quietly. Fill a gas jar full, of water, cover it with a greased glass plate, *G*; turn the jar and cover upside down, and remove the plate while the mouth of the gas jar is below the surface of the water in the gas trough. Place the mouth of the jar over the hole in the floor of the "beehive." When the jar is full of gas, close the mouth of the jar with the glass plate and remove the vessel from the collecting trough. Stand the jar mouth downwards, and collect several jars of gas in a similar way.

Properties.—Plunge a lighted taper into a jar of the gas held mouth downwards; the gas is combustible, for it burns with a scarcely visible blue flame at the mouth of the jar, and the taper is extinguished, showing that the gas is a non-supporter of combustion. The gas can be poured upwards from one jar to another as illustrated in Fig. 11, and it can be proved that the gas has actually been transferred from the one vessel to

¹ The tubulated bottles for washing gases appear to have been first described by Peter Woulfe in 1784, hence the term Woulfe's bottles, not "Woulff's bottles."

² The discovery of the water pneumatic trough is often attributed to Stephen Hales, about 1730; J. Priestley afterwards used mercury in place of water, and this enabled him to manipulate gases soluble in water.

the other by testing the contents of each jar with a lighted taper. The gas is therefore lighter than air; indeed, for many purposes there is no need to use the pneumatic trough for collecting hydrogen. Bring the gas jar mouth downwards over a jet of hydrogen. The hydrogen collects at the top of the jar, and displaces the air downwards—hence the term *collecting gases by the downward*¹ *displacement of air*. Later on we shall meet another important gas which turns clear lime water turbid—hydrogen does not. As indicated above, a mixture of hydrogen with oxygen or air is violently explosive. This can be illustrated by mixing two volumes of hydrogen gas with either one volume of oxygen or five volumes of air in a soda-water bottle. A lighted taper applied to the mouth of the bottle causes the gas to detonate violently. The combustion of the whole mass is almost instantaneous. The explosion is so violent that N. Lemery believed thunder to be due to the “fulminations of hydrogen.”

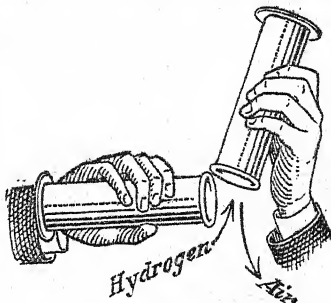


FIG. 11.—Pouring Hydrogen upwards.

These experiments have taught us that hydrogen is a colourless gas without taste or smell. Its specific gravity is very low compared with air, in other words, hydrogen gas is much lighter than air. The gas is inflammable, and burns in air with an almost colourless flame, while a lighted taper is extinguished when plunged into the gas. Hydrogen gas has no apparent action on clear lime water.

§ 2. Burning Hydrogen in Air.

Joseph Priestley has told us, in 1776, that his friend, J. Warltire, noticed that when a flame of hydrogen is allowed to burn in air confined under a bell-jar, the whole of the receiver appeared to be filled with a white powdery substance, and when the flame goes out, the air left in the jar was perfectly noxious. In the same year P. J. Macquer inquired whether the flame of hydrogen evolves soot or smoke, and he thus described his experiment:

By placing a saucer of white porcelain in a jet of inflammable gas burning tranquilly at an orifice, I found that the part of the saucer which the flame licked was moistened by small drops of liquid as clear as water, and which, in fact, appeared to be nothing but pure water.

It is probable that J. Warltire's white cloud was nothing but condensed steam. P. J. Macquer has been blamed because he felt no astonishment at that which is really astonishing—the appearance of the water—and because he did not stop to inquire: Whence came the water? He did not see a great discovery begging for recognition. Hence, asked F. J. Arago,

¹ Many writers call this *collecting the gas by upward displacement*. The student must therefore make perfectly clear there is no confusion in his own use of the term.

Is genius in the observational sciences to be reduced to the faculty of asking an appropriate why? The inquiry can be made (1) What happens to the surrounding air during the burning of a jet of hydrogen? and (2) Is the product of the action really water?

J. Warltire's experiment can be modified in the following way: Fill a gas holder,¹ *M*, Fig. 12, with water. Close the two stopcocks, remove the stopper *N* and place the delivery tube from the hydrogen apparatus (Fig. 10) in the opening. Water is displaced² by the hydrogen. When full, close *N* with the stopper. By keeping the large funnel, *F*, filled with water, and regulating the two stopcocks, *S* and *S*₁, hydrogen can be displaced from the gas holder at any required speed.³ Connect a glass tube, bent as shown in the diagram, with the exit tube from the gas holder, interposing a calcium chloride tower *G*, packed with granulated calcium chloride, and a plug of glass wool at each end. The object of this tube is to remove moisture from the gas.⁴ Place a bell-jar, *K*, in a dish of distilled water, *H*, Fig. 12, and arrange the delivery tube *J* so that it can be quickly

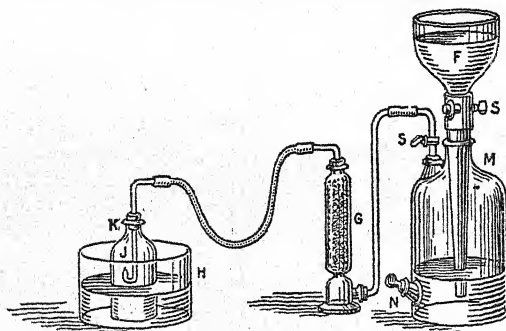


FIG. 12.—The Action of burning Hydrogen in Air.

fixed as illustrated in the diagram. Collect a test tube of the gas by downward displacement as it leaves the exit tube in order to make sure that all air has been displaced from the tubes.

When the test is satisfactory, light the jet of hydrogen, and adjust the flame until it is about the size of a candle-flame; lower the burning jet into the cylinder of air as shown in Fig. 12. Note that the water rises in the jar, and that the flame of hydrogen gradually expires. Immediately this occurs, stop the current of hydrogen by means of the stopcock, otherwise hydrogen gas will pass from the gas holder and mix with the residual air.

The gas remaining in the jar has quite similar properties to the gas remaining after mercury is calcined in air, p. 23. Consequently, it is inferred that when hydrogen burns in air, it unites with the oxygen and leaves nitrogen behind. If the experiment be carefully done, four-

¹ There is no special need for a gas holder in this particular experiment. Kipp's apparatus may be used, as illustrated in 4, Fig. 13. The gas holder is introduced here to illustrate one of the methods of storing gases in the laboratory.

² This is done over a sink to carry off the surplus water.

³ When hydrogen is to be burned, a jet of hard glass should be used, or, better, a tip of platinum or a piece of quartz glass tubing should be attached by a piece of rubber tubing to the delivery tube. The hydrogen flame is very hot and soon melts ordinary soft glass.

⁴ Concentrated sulphuric acid is also frequently used to absorb moisture from gases. Phosphorus pentoxide is used in special cases where a very powerful absorbent is needed for very small amounts of moisture. The drying agent used must not react with the gas.

fifths of the original volume of air remain. The burning hydrogen removes one-fifth of the original volume of air. Hydrogen does not burn in the residual nitrogen. A certain amount of "dew" collects on the inner walls of the bell-jar, but that, of course, may come from the water in the dish below. In fine, we have reasons for supposing that hydrogen, in burning, combines with oxygen to form an oxide of hydrogen in the same sense that mercury, when calcined in air, combines with oxygen to form mercuric oxide. It remains to try to isolate the hydrogen oxide whose existence we have just inferred, but not proved.

P. J. Macquer's experiment can be modified in the following manner: Fit up the apparatus previously described for the preparation of hydrogen, Fig. 12, or use one of the numerous modifications of Kipp's apparatus, say, A, Fig. 13. Kipp's apparatus is very convenient when a steady current of hydrogen is needed for some time; or when variable quantities of gas are required intermittently.¹ The Kipp's apparatus may conveniently be used instead of the gas holder in the preceding experiment. Connect the Kipp's apparatus with the drying tower packed with calcium chloride, B, and fitted with a jet, C, for burning the hydrogen. A piece of narrow quartz-glass tubing makes an excellent jet. If this is not available, use a jet of hard "combustion" glass tubing. Test the gas, to ensure the absence of air, by bringing a *dry* test

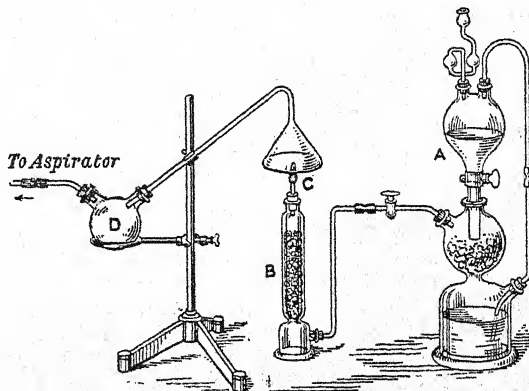


FIG. 13.—The Synthesis of Water.

tube over the jet of gas issuing from the drying tower. Bring a lighted taper to the mouth of the tube. The hydrogen should burn quietly; in that case, it will be noticed that a kind of "dew" collects on the inside of the test tube. Bring the jet of burning gas under the inverted funnel fitted, as shown in the diagram, with a bulb, D, connected with an aspirator for sucking a gentle current of air through the bulb. The current of air carries along the products of combustion from the hydrogen flame. A clear colourless liquid collects in the bulb.² This

¹ Zinc is placed in the middle bulb, and dilute sulphuric acid (1 vol. acid; 8 vols. water) or hydrochloric acid in the lower and upper bulb, as shown in the diagram. Open the stopcocks until the gas has displaced the air. The velocity of the current of gas is regulated by the stopcocks. When the stopcock is closed, the acid is forced away from the zinc. The side tube helps to prevent an accumulation of spent acid near the zinc. There are over a hundred modifications of the principle for supplying a continuous stream of gas.

² The bulb can be cooled, if desired, by resting it in a funnel, and allowing water to run over the bulb and out from the bottom of the funnel to the sink.

liquid has all the properties of water—clear, colourless, tasteless, no smell, freezes at 0° , boils at 100° , etc. It is therefore provisionally inferred that water is a hydrogen oxide formed when hydrogen burns in air.

§ 3. Morley's Experiment on the Composition of Water by Weight.

In the determination of atomic weights, a small number of values are to be regarded as fundamental. They are the standards of reference; and by comparison with them all the other atomic weights are established. The atomic weights of hydrogen and oxygen are primary; that is, one or other of them is the basis of the entire system of atomic weights.—F. W. CLARKE.

E. W. Morley (1895) has a very fine application of the principle underlying this experiment. Known weights of pure dry hydrogen and pure

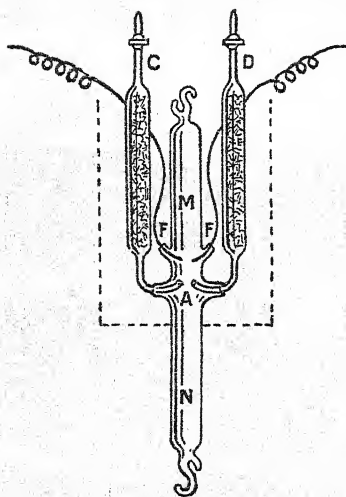


FIG. 14.—Morley's Experiment.—
Synthesis of Water.

dry oxygen were stored in two large glass globes.¹ The globe containing oxygen was connected with *C*, Fig. 14. The oxygen passed through a layer of phosphorus pentoxide,² and thence into the glass chamber *M* *via* one of the jets, *A*; the globe containing hydrogen was similarly connected with another tube, *D*, containing phosphorus pentoxide, and the hydrogen led into the chamber *M* *via* one of the jets *A*. The rates at which the gases enter the chamber was regulated by suitable stopcocks. The chamber *M* was previously evacuated and weighed. One of the gases, say oxygen, was allowed to enter *M*, and electric sparks were passed across the terminals *F* just over the jets *A*. Hydrogen was led into the apparatus and ignited by the sparks. The rates at which hydrogen and oxygen entered the chamber were regulated so that the formation of

¹ The hydrogen was prepared by heating palladium hydride, and the oxygen by heating potassium chlorate. The apparatus for storing and drying the hydrogen and oxygen is not shown in Fig. 14.

² The phosphorus pentoxide is not intended to dry the entering gases—these have already been dried.

The phosphorus pentoxide tubes prevented the escape of water vapour. The amount of water formed was determined from the difference in the weights of the system *M* before and after the experiment. The amounts of hydrogen and oxygen used were determined from the weights of the corresponding globes before and after the experiment. The amount of water formed was determined from the increase in the weight of the above-described vessel before and after the combustion. Morley, as a mean of eleven experiments, found that:

Hydrogen used	3.7198 grams.
Oxygen used	29.5335 grams.
Water formed	33.2530 grams.

Hence, taking oxygen = 16 as the unit for combining weight, it follows that *16 parts by weight of oxygen combine with 2.016 parts by weight of hydrogen to form 18.016 parts of water*—within the limits of the small experimental error.

§ 4. Dumas' Experiment on the Composition of Water by Weight.

After a careful examination of all analytical researches made for the determination of atomic weights, I emphatically declare that the researches of Dumas are the most important of all, marking as they do the beginning of analysis of precision, and offering also the first instance of a true series of determinations such as is required to furnish the absolute value of the atomic weights.—G. D. HINRICHS.

Hydrogen does not combine easily with many of the elements, but it readily combines with oxygen, chlorine, fluorine, lithium, and a few others. Hydrogen will very often remove oxygen and chlorine from their compounds with the other elements. Thus, when hydrogen is passed over hot ferric oxide, lead oxide, nickel oxide, copper oxide, etc., the hydrogen combines with the oxygen of the oxide and leaves behind the metal. In these experiments, the hydrogen is said to be *oxidized*; and the metallic oxide, *reduced*, or *deoxidized*.

If a known amount of copper oxide be employed, and the water formed be collected and weighed, the weight of the reduced copper oxide will show how much oxygen has been used in forming a definite amount of water. This was done by J. B. A. Dumas in 1843. His experiment is not the best of its kind, although it was the best of its time, and it has long held an honoured place in chemical text-books. The experiment illustrates some important principles, and it is therefore here described in outline.

The preparation of pure hydrogen.—The hydrogen was prepared by the action of zinc on sulphuric acid. It might be thought that pure zinc and pure sulphuric acid should be used. Experiment shows, curiously enough, that the action is so very, very slow, that it is often stated that "absolutely pure sulphuric acid, even when diluted with pure water, has no action on perfectly pure zinc." Moreover, it is exceedingly difficult to prepare pure zinc and pure sulphuric acid. Hence, pure reagents were not used for the preparation of the hydrogen. Accordingly, the gas may contain nitrogen and oxygen derived from the air; sulphur dioxide and hydrogen sulphide derived from the reduction of the sulphuric acid by the hydrogen, carbon dioxide, arsenic hydride (if the acid or the zinc contained

arsenic); hydrogen phosphide (if the zinc or the acid contained phosphorus); nitrogen oxides (if the acid contained nitrogen oxides); and water vapour. Accordingly, Dumas (1843) used sulphuric acid, which had been well boiled to get rid of dissolved air, and then passed the hydrogen through a series of U-tubes—Fig. 15—containing: (1) pieces of glass moistened with lead nitrate to remove hydrogen sulphide; (2) solution of silver sulphate to remove arsenic and phosphorus compounds; (3) solid potassium hydroxide to remove sulphur dioxide, carbon dioxide, and nitrogen oxides; ¹ and (4) phosphorus pentoxide to remove moisture ¹ not absorbed by the solid potassium hydroxide. The phosphorus pentoxide tubes were placed in a freezing mixture. The tube marked (5) in the diagram contained phosphorus pentoxide. It was weighed before and after the experiment. If no change in weight occurred, it was assumed that the hydrogen passing through was quite dry.

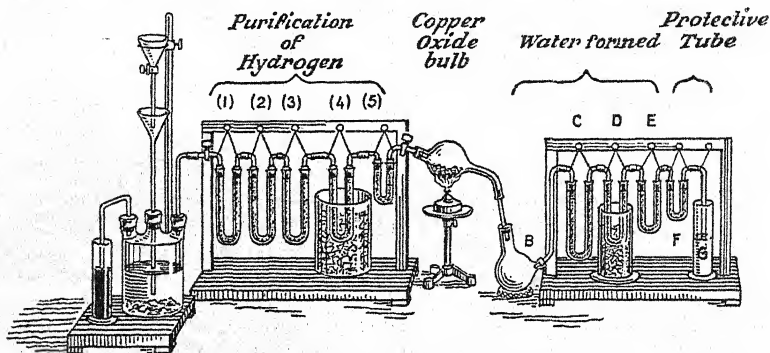


FIG. 15.—Dumas' Experiment (abbreviated).

The experiment.—The purified hydrogen was passed through a weighed bulb, A, containing copper oxide, and heated by the spirit lamp underneath. Most of the water condensed in the bulb B, and the remainder was absorbed in the U-tube C containing solid potassium hydroxide, and in D and E containing phosphorus pentoxide. The phosphorus pentoxide tube D was kept cool by a freezing mixture. The three tubes C, D, E, and the bulb B, were weighed before and after the experiment. The last U-tube, F, containing phosphorus pentoxide was followed by a cylinder, G, of sulphuric acid through which hydrogen escaped. The vessels F and G were not weighed; they served to protect the other tubes from the external atmosphere.

The results.—The average of nineteen experiments by Dumas (1843) gave:

Copper oxide lost in weight	44.22 grams.
Water produced	49.76 grams.
Hydrogen (by difference)	5.54 grams.

Hence, every 16 parts by weight of oxygen combined with 2.004 parts

¹ Dumas used three potassium hydroxide tubes, and two phosphorus pentoxide tubes. I have taken the liberty of showing only one of each in the diagram.

by weight of hydrogen to form water. The latter determination of Morley gave 16:2.016. There is a curious error in Dumas' experiment which, if not corrected, makes the result a little high. The reduced copper retains some hydrogen very tenaciously (see occlusion of hydrogen by the metals, p. 125). Other objections to Dumas' experiment have been made: (1) The expulsion of the air from the large apparatus is difficult; (2) The air absorbed by the sulphuric acid is slowly evolved along with the hydrogen when the acid acts on zinc; (3) The copper oxide is contaminated with nitrogen absorbed from the air; (4) The slight reduction of sulphuric acid by hydrogen, forming gaseous sulphur dioxide; and (5) The complete drying of the hydrogen is very difficult. In approximate work we may take it that 2 parts by weight of hydrogen combine with 16 parts by weight of oxygen to form 18 parts of water.

§ 5. The Decomposition of Water by Metals.

Water remains permanent and stable so long as the balance of forces between its constituent elements is maintained, but in the presence of another body which can unite with one of these elements the water may be decomposed; one element—say, hydrogen—is set free and the other unites with the agent of destruction to form a new compound—an oxide.—ANON.

In the preceding experiments, water has been synthesized from its elements. Let us now examine some methods of decomposing water into

its elements — analysis.

Fill an iron, porcelain, or hard glass tube—60 cm. long and 1.5 cm. diameter—with bright iron turnings or bright iron nails. In Fig. 16 a hard glass tube is used. This is

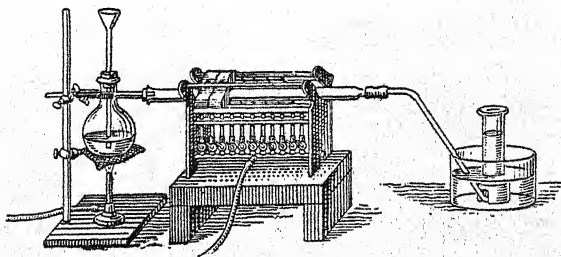


FIG. 16.—Decomposition of Steam by Hot Iron.

drawn out at one end as shown in the diagram. This end is fitted with a delivery tube dipping in a gas trough. A roll of previously ignited asbestos paper, 6 cm. long, is inserted in the opposite end. This end is closed with a red rubber stopper and the exit tube of the flask so arranged that it passes a short distance into the core of the asbestos paper. The asbestos roll, later on, prevents the liquid water coming into contact with the hot glass and breaking the tube. Water is boiled in the flask, and the steam, passing through the red-hot iron turnings, is decomposed. When all the air has been driven out of the apparatus, hydrogen may be collected in the gas jar. The usual tests for hydrogen, indicated on p. 60, may be applied.

Lavoisier made a similar experiment to this in 1783, and stated that the metallic iron "is converted into a black oxide precisely similar to that produced by the combustion of iron in oxygen gas." The iron is oxidized by the water, and the water is reduced, forming "a peculiar inflammable

gas," which Lavoisier named "hydrogen," because "no other term seemed more appropriate." The word signifies the "generative principle of water," from the Greek ὑδωρ (hydor), water, and γεννᾶω (gennao), I generate or produce. The German word for hydrogen is "Wasserstoff."

If zinc dust be used in place of iron, the temperature need not be much higher than the boiling point of water, since zinc reduces steam and forms zinc oxide at a comparatively low temperature. If a strip of magnesium ribbon be placed in a bulb of a hard glass tube and heated in a current of steam at a red heat, the metal appears to burst into flame, forming magnesium oxide. The resulting hydrogen can be ignited if the jet of steam be not too vigorous. Metallic calcium decomposes cold water and gives off hydrogen, but the action slows down very soon, probably because the calcium hydroxide is not all dissolved by the water, and in consequence a crust of this substance forms over the surface of the metal. The calcium can be advantageously warmed with water in a flask, connected directly with a delivery tube leading to the gas trough. If the water is not free from carbonates, a crust of calcium carbonate also forms

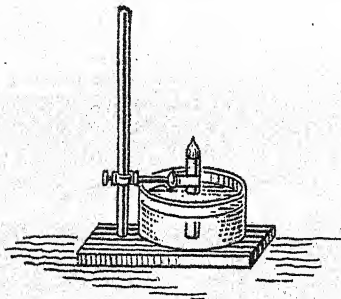


FIG. 17.—The Action of Sodium on Water.

over the surface of the metal. Calcium hydroxide is formed as well as hydrogen. Sodium decomposes cold water, giving off hydrogen, and forming sodium hydroxide. The experiment is liable to unpleasant explosions when the sodium is confined so as to enable the resulting hydrogen to be collected. The following is an easy way of showing the action.¹ A glass tube—4 or 5 cm. long and 1.5 cm. diameter, and open at both ends—is thoroughly dried inside. This is lowered and clamped vertically in a dish of water so as not to wet the sides of the tube above the level of the water, Fig. 17. A small piece of dry sodium, about 2 mm. in diameter, is dropped into the tube. The hydrogen evolved can be lighted at the upper end of the tube. Sodium amalgam—that is, a solution of metallic sodium in mercury—decomposes water much less turbulently than sodium alone. The result is similar when a small piece of potassium—3 or 4 mm. diameter—is placed on water, but it reacts so violently that the temperature rises high enough to set fire to the hydrogen.² This burns with a flame tinged violet, owing to the presence of the vapour of potassium; the hydrogen produced by the action of sodium on water burns with a yellow flame, owing to the contamination of the hydrogen by the vapour of sodium.

This set of experiments gives a series of metals which appear to react with water with increasing violence; the metals seem to have an increasing "avidity" or "affinity" for oxygen:

Iron, zinc, magnesium, calcium, sodium, potassium.

¹ If there is an explosion, no particular harm is done.

² A sheet of plate glass should be held between the metal on the water and the operator.

§ 6. The Decomposition of Water by Electricity.

Electricity is the key which will open a way into the innermost parts of nature.—J. W. RITTER.

W. Nicholson and A. Carlisle, May 2, 1800,¹ happened to put a drop of water in contact with two wires from an electric battery and noticed the formation of small bubbles of gas about the tips of the wires when the tips of the wires were not in contact. They then immersed the two wires in a glass of water, and found that gases were formed about both wires. They found the gas collected at one wire to be hydrogen, and at the other wire, oxygen. Two volumes of hydrogen were collected for every volume of oxygen. The gases were mixed and exploded. The result was water. This is very interesting. We have seen that chemical combination can produce an electric current; here an electric current is used to produce chemical decomposition.

The electrolysis of water.—It will be convenient to modify Nicholson and Carlisle's experiment. A trough, Fig. 18, is half filled with water

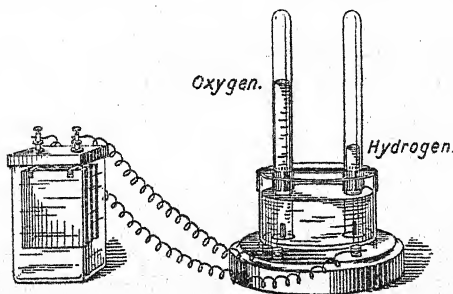


FIG. 18.—Electrolysis of Water—Gases separated.



FIG. 19.—Electrolytic Gas.

slightly acidulated with sulphuric acid. Test-tubes full of acidulated water are placed in the position shown in the diagram over two plates of gold or platinum. The plates are put in communication with an accumulator or galvanic battery.² During the passing of the electric current, bubbles of gas from about the metal plates rise into the test-tubes. More gas is given off at one plate than the other. The gas in each tube can be examined by means of a lighted taper or otherwise. In the one tube, the taper burns with the "blinding brilliance" characteristic of oxygen; and the gas in the other tube burns with the blue flame characteristic of hydrogen. Some of the water has disappeared, but no change can be detected in the amount of sulphuric acid mixed with the water. Hence it is inferred that the water, not the acid, has been decomposed. The experiment succeeds equally well if a very dilute solution of sodium or potassium hydroxide be used with nickel or iron electrodes. Here again the water, not the alkali, is decomposed. These reagents are used because

¹ J. W. Ritter noticed the decomposition of water by an electric current a year earlier—1799.

² There are scores of different ways of doing this experiment. Some of the instruments are very ingenious, but easily broken.

water alone does not conduct an electric current very well. In fact, pure water is said to be a non-conductor of electricity. Dilute solutions of acids or alkalis are good conductors. If iron electrodes are used in the acidulated solution, much of the oxygen formed during the decomposition of the water is used in oxidizing the metal.

A mixture of one volume of oxygen and two volumes of hydrogen, called electrolytic gas or detonating gas, is often wanted in gas analysis, etc. This is easily provided by placing both electrodes under one receiver. The apparatus illustrated in Fig. 19 is often used for this work. Electrolytic oxygen contains a little ozone and hydrogen peroxide if prepared by the electrolysis of acidulated water, but not if a solution of barium hydroxide be electrolysed.

The formula for water used to be written HO when the atomic weight of hydrogen was taken unity, and oxygen 8. This agrees quite well with the determinations of Morley and of Dumas. But we naturally ask for an explanation of the result of the electrolysis of water. Does an atom of hydrogen occupy twice the volume of an atom of oxygen? This subject will be taken up in the next chapter.

§ 7. Cavendish's Experiment on the Synthesis of Water by Volume.

It is curious to note the changing fortunes of water in the history of chemistry. First the matrix of the whole universe; then only one of the four elements though the chief of the quaternion; and at last discovered to be nothing but a liquid product of combustion, one oxide among many, the mere ash, rust, or calx of so much burnt hydrogen.—S. BROWN.

The older chemists considered water to be an element. They were quite right so far as their knowledge went, p. 25, because they did not know how to decompose it into simpler substances. In the spring of 1781, J. Priestley made what he called "a random experiment" to "entertain a few philosophical friends," in which a mixture of "inflammable air" with oxygen or atmospheric air was exploded in a closed vessel by means of an electric spark. The sides of the glass vessel were found "bedewed" with moisture after the explosion, but Priestley paid no particular attention to the phenomenon. H. C. Cavendish looked upon the deposition of the dew as a fact "well worth examining more closely," and immediately followed up the subject, in 1781, by exploding a mixture of "one measure of oxygen with two measures of inflammable air (hydrogen)" in a closed vessel. No gas remained in the globe after the explosion, but the hydrogen and oxygen lost their gaseous form, and produced a certain weight of water. The vessel and its contents underwent no change in weight, or parted with anything ponderable during the explosion, whilst a certain volume of gas was replaced by a certain weight of water. Hence Cavendish deduced that liquid water consists, weight for weight, of the hydrogen and oxygen gases lost in its production. Cavendish's results were communicated to Priestley and to Lavoisier not later than the summer of 1783, and published in 1784.¹

The experiment can be illustrated in the following manner: A stout

¹ James Watt, of engineering fame, expressed the *opinion*, in 1783, that Priestley's experiment meant that "water is composed of dephlogisticated air (oxygen) and inflammable air." There was much confusion about that time in the use of the term "inflammable air," and it is by no means clear that Watt meant by "inflammable air" what we understand by "hydrogen" to-day.

glass vessel, *A*, is fitted with a stopcock, *C*, at one end, and with a piece of strong pressure tubing, *D*, connected with a reservoir at the other end, Fig. 20. A pair of platinum wires, *T*, are sealed into the stout glass measuring vessel just below the stopcock.

These wires are put in communication with an induction coil, which in turn is connected with an accumulator. The tube *A* is called the eudiometer, or the explosion tube. This is filled with mercury by adjusting the levelling tube *B* and the stopcock *C*. A mixture containing one volume of oxygen and two volumes of hydrogen is introduced into the explosion tube *via* the stopcock *C* and by depressing the levelling tube. When the explosion tube is about half or three-fourths filled, read the volume of its contents by bringing the mercury to the same level in both levelling tube and explosion tube. Then depress the levelling tube so that the mercury falls nearly to the bottom of the explosion tube. Pass a spark from the induction coil through the wire terminals of the explosion tube. The gases explode, and the level of the mercury is again adjusted after the apparatus has stood for a few minutes in order to regain the temperature of the room. The mercury rises nearly ¹ to the level of the stopcock.²

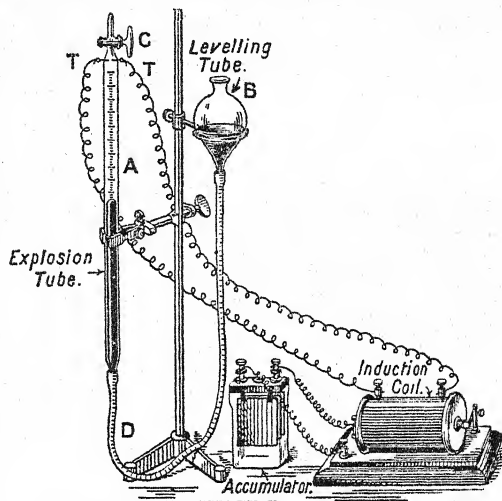


FIG. 20.—Synthesis of Liquid Water by Volume.

Suppose the experiment be repeated a number of times with, say, one volume of oxygen and three volumes of hydrogen—one volume of hydrogen remains after the explosion; again try the experiment with two volumes of oxygen and two volumes of hydrogen—one volume of oxygen will remain uncombined after the explosion. It is inferred, from this experiment, that two volumes of hydrogen and one volume of oxygen combine to form water, and if an excess of either oxygen or hydrogen be present, the excess will remain uncombined after the reaction.

Gas analysis.—If a known volume of a gas containing hydrogen be mixed with an excess of air or oxygen; or if a known volume of a gas containing oxygen be mixed with an excess of hydrogen and exploded in a eudiometer, the contraction which occurs represents the volume of water

¹ The mixed gas probably contained a little air, and probably a slight excess of either oxygen or hydrogen.

² The advantage of the forms of explosion vessel, Figs. 20 and 21, lies in the fact that the explosion takes place under diminished pressure, and is not so liable to fracture the apparatus, because it is less violent.

formed, and the corresponding amount of the gas under investigation can be determined. For example, 20 c.c. of air was mixed with 20 c.c. of hydrogen and exploded. The mixed gases, after the explosion, occupied 28 c.c. Hence, the contraction was 12 c.c. Hence, 12 c.c. of the mixture combined to form water. Of this two-thirds must have been hydrogen and one-third oxygen. Hence, the original 20 c.c. of air contained 4 c.c. (*i.e.* one-third of 12 c.c.) of oxygen. This illustrates an important principle used in gas analysis.

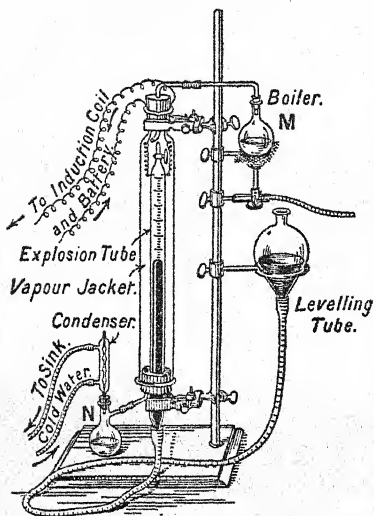
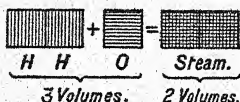


FIG. 21.—Synthesis of Steam by Volume.

110°, or amyl alcohol, boiling at about 130°. The lower end of the jacket is connected with a flask and condenser *N*, so that the amyl alcohol can be recovered. When the amyl alcohol is steadily boiling, and the explosion tube has been filled as described in the preceding experiment, the gases are sparked. In a few minutes, when the temperature has had time to adjust itself, bring the levelling tube in position for a reading. It will be found that the steam occupies just two-thirds the original volume of the mixed gases. Otherwise expressed,



Hence, it is inferred that when water is synthesized at a temperature above the point of condensation—100°—two volumes of hydrogen react with one volume of oxygen to form two volumes of steam. It is necessary to correlate the different results described in this chapter when water is synthesized by volume and by weight.

Questions.

1. 50 c.c. of oxygen are mixed with 500 c.c. of hydrogen, both are measured at normal temperature and pressure, and an electric spark is passed through the

mixture ; what volume, if any, of gas will remain, and how would you ascertain what it is ?—*Science and Art Dept.* Hint : 50 c.c. of oxygen unite with 100 c.c. of hydrogen to form water, and most of the water condenses to a liquid ; 400 c.c. of moist hydrogen remain.

2. R. Bunsen (1846) mixed 436.97 c.c. of dry air with hydrogen. The mixed gases, occupying 672.74, were sparked, and the residual gas, when dried, occupied 403.88 c.c. (a) What is the percentage composition of dry air, assuming the nitrogen of the air is not affected by the explosion ? (b) Was sufficient hydrogen added to combine with all the oxygen of the air ? Ansr. (a) 20.91 per cent. of oxygen and 79.09 per cent. of nitrogen. (b) Yes. An excess of 58.20 c.c. of hydrogen was added.

3. A mixture of 5 volumes of hydrogen and 3 volumes of air is surrounded by a hot jacket at a constant temperature of 110° . What change of volume will occur after the mixture has been sparked, assuming that air contains 21 volumes of oxygen and 79 volumes of inert nitrogen ? Here 3 vols. of air contain 0.63 vol. of oxygen, and 2.37 vols. of nitrogen ; 0.63 vol. of oxygen unites with 1.26 vols. of hydrogen to form 1.26 vols. of steam. Hence, the composition of the mixture after sparking will be 2.37 vols. of nitrogen, 1.26 vols. of steam, and 3.74 vols. of hydrogen. Total, 7.37 vols. The volume of the original mixture is 8 vols. Hence a contraction of 0.63 vol. occurs. This question could have been more simply answered by noting that the hydrogen simply removes 0.63 vol. of oxygen from the mixture.

CHAPTER IV

COMBINATION BY VOLUME

§ 1. Gay-Lussac's Law of Combining Volumes.

Thou hast ordered all things in measure, and number, and weight.—WISDOM OF SOLOMON.

NOT very long after Dalton had directed the attention of chemists to the relations subsisting between the weights of bodies which combine in different proportions, Gay-Lussac established a similar correspondence between volumes of combining gases. A. von Humboldt, the naturalist and explorer, collected samples of air from different parts of the world, and, with the aid of J. F. Gay-Lussac, analysed the different samples with the idea of finding if the composition of air was variable or constant. Gay-Lussac used Cavendish's process—explosion of a mixture of air and hydrogen gas (p. 68). As a preliminary, Humboldt and Gay-Lussac investigated the proportion by volume in which hydrogen and oxygen combine, and found the ratio of hydrogen to oxygen, by volume, to be nearly as 2 : 1. If either hydrogen or oxygen was in excess of these proportions, the excess remained, after the explosion, as a residual gas. Humboldt and Gay-Lussac (1805) found :

Vols. oxygen.	Vols. hydrogen.	Vols. residue.
100	300	101.1 hydrogen
200	200	101.7 oxygen

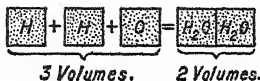
After making corrections for impurities, etc., in the gases, Gay-Lussac and Humboldt stated that "100 volumes of oxygen required for complete saturation 199.89 volumes of hydrogen, for which 200 may be put without error." A. Scott (1893) found, as the result of twelve experiments on the volumetric composition of water, that oxygen and hydrogen combine very nearly in the ratio 1 : 2.00245 by volume.

Struck by the simplicity of the relation thus found, J. F. Gay-Lussac (1808) followed up the subject by numerous experiments with different gases. As a result, he concluded that "gases always combine in the simplest proportions by volume." For instance, one volume of hydrogen combines with one volume of chlorine forming two volumes of hydrogen chloride; this fact can be represented diagrammatically :

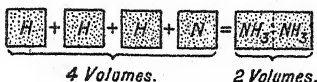


Two volumes of hydrogen combine with one volume of oxygen forming

two volumes of water vapour (which condenses to liquid water if the temperature be below 100°).



Three volumes of hydrogen and one volume of nitrogen form two volumes of ammonia. Thus:



There are slight deviations with the gases which show deviations from the laws of Boyle and Charles, but the experimental results are such as to leave no doubt that Gay-Lussac's generalization is valid, and accordingly, we define Gay-Lussac's law; when gases react together, they do so in volumes which bear a simple ratio to one another, and to the volume of the gaseous product of the action. It is assumed, of course, that the initial and final products of the reaction are under the same conditions of temperature and pressure.

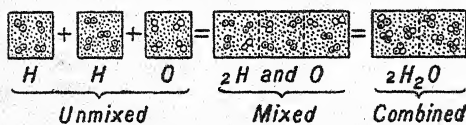
We traced the remarkable way in which elements combine by weight to a peculiarity in the constitution of matter; so here, we are tempted to make a similar quest. It follows at once (1) if elements in a gaseous state unite in simple proportions by volume, and (2) if the elements also unite in simple proportions by atoms, then the number of atoms in equal volumes of the reacting gases must be simply related. With J. Dalton let us make a guess. Assume that equal volumes of the different gases under the same physical conditions contain an equal number—say n —of atoms. Then, when two volumes of hydrogen react with one volume of oxygen to form two volumes of steam, we have $2n$ atoms of hydrogen reacting with n atoms of oxygen to form $2n$ "compound atoms" of steam. Hence, two atoms of hydrogen react with one atom of oxygen to form two "compound atoms" of steam. In that case, every atom of oxygen must be split into half an atom to make two "compound atoms" of steam. This contradicts the fundamental postulate of the atomic theory, or Dalton's aphorism: "Thou knowest no man can split an atom," meaning, of course, that atoms are indivisible in chemical reactions. Similar contradictions are encountered in nearly every case of combination between gases, hence Dalton claimed this guess to be untenable; we must try another. There is such a marked uniformity in the deportment of elementary and compound gases with respect to variations of temperature and pressure, that it is not very probable any essential difference will be found in the constitution of elementary and compound gases.

§ 2. Avogadro's Hypothesis.

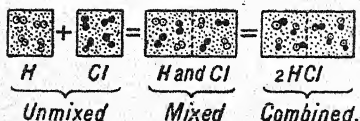
Advances in knowledge are not commonly made without the previous exercise of some boldness and licence in guessing.—W. WHEWELL.

A. Avogadro (1811) pointed out that the fallacy in Gay-Lussac's reasoning can be avoided if we distinguish clearly between elementary

atoms and the small particles of a gas. Assume that the small particles of a gas are aggregates of a definite number of atoms. Avogadro called these aggregates *molecules* in order to distinguish them from the ultimate atoms. The term "molecule" is the diminutive form of the Latin word *moles*, a mass. Each molecule of an elementary gas contains the same number and kind of atoms. For the sake of simplicity, assume that each molecule of hydrogen gas is composed of two atoms of hydrogen, and make a similar assumption for oxygen gas. Hence, modify Gay-Lussac's guess and assume that equal volumes of all gases contain the same number of molecules. Suppose that two volumes of hydrogen contain $2n$ molecules of hydrogen, then one volume of oxygen will contain n molecules. These react to form $2n$ molecules of steam—each molecule of steam contains two atoms of hydrogen and one atom of oxygen. Hence, as W. K. Clifford expressed it, although atoms cannot be split so that one atom of oxygen enters into the composition of two molecules of water, yet one molecule of oxygen can be divided between two molecules of water. You cannot put 50 living horses into 100 stables so that there will be the same amount of horse in each stable, but you can divide 50 *pairs* of horses among 100 stables. The idea can be more clearly illustrated by means of the subjoined diagrams analogous to those used by A. Gaudin about 1832. Each square represents one volume of a gas. Each volume contains n molecules. We do not know the numerical value of n , but, for the sake of simplicity, take $n=4$. It makes no difference to the final conclusion what numerical value we assign to n . Then we have:



Again, with hydrogen and chlorine,



Diagrams similar to these were used by A. Gaudin, 1832. It must not be supposed for one moment that these diagrams are intended as pictures of the actual molecules. They are to be regarded as aids to the understanding of how Avogadro's hypothesis has led chemists to conclude that the molecules of gaseous elements are really compounded atoms, and how Avogadro's hypothesis reconciles the observed volume relations during the combination of gases with the atomic theory.

We have assumed for the sake of simplicity, that the molecule of water contains three atoms, and that each molecule of hydrogen and oxygen contains two atoms. As a matter of fact, all we can infer from the observed facts is that the molecule of oxygen is split into halves, and in the absence of evidence to the contrary, we must assume for every substance the simplest molecular structure consistent with the observed facts.

Avogadro thus modified the atomic hypothesis and introduced the

conception of two orders of minute particles, (1) the atom of the unit of chemical exchange; and (2) the molecule or the smallest particle of an element or compound which exists free in a gas. This definition of a molecule is usually extended into the less satisfactory definition: A molecule is the smallest particle of an element or compound which exists in a free state; meaning that the specific properties of a substance depend on the component particles remaining intact. If the molecules be subdivided (or augmented) the substance will no longer have the same specific properties, because the nature of the component particles (molecules) is different. A diatomic molecule for gaseous chlorine, hydrogen, and oxygen at ordinary temperatures, is a satisfactory explanation of what we know to-day, but it is possible that at some future date the evidence will compel us to consider these molecules to be tetra- or hexatomic. This will not materially affect the principle as indicated above. The molecule of mercury, as we shall see later, is supposed to be monatomic, and the molecule of sulphur hexatomic. Avogadro's hypothesis states that equal volumes of all gases, at the same temperature and pressure, contain the same number of molecules. A. M. Ampère tried unsuccessfully to extend the same hypothesis to solids in 1814.

Increasing knowledge has made the hypothesis more and more probable; it has been tested in hundreds of experiments, and never found wanting. The hypothesis has done such good service in giving a rational explanation of many different phenomena that it has been accepted as a fundamental truth.¹ Avogadro's hypothesis is the basis of the current theory of chemistry.

§ 3. The Relative Weights of the Molecules.

In order to bring into harmony all the branches of chemistry, we must have recourse to the complete application of the theory of Avogadro and Ampère in order to compare the weights and the numbers of the molecules.—S. CANNIZZARO.

Avogadro's hypothesis has proved to be one of the most suggestive and fruitful hypotheses in the development of chemistry. It has correlated what appeared antagonistic and contradictory; it has harmonized what seemed discordant and confused, and made Dalton's atomic hypothesis a clear, intelligible, and fertile theory. Had it not been for this development, Dalton's hypothesis was in a fair way of being "sentenced to sterility and oblivion" (A. Würtz, 1887).

By definition, the relative density of a gas is a number which represents how much heavier any volume of the gas is than an equal volume of the standard gas—generally hydrogen—measured at the same temperature and pressure—generally at 0° and 760 mm. pressure. Thus, the relative density of steam is 8.95. This means that any volume, say a litre of steam, is nearly nine times as heavy as the same volume of hydrogen.²

¹ Deviations from Avogadro's law are observed with those gases which deviate from Boyle's and Charles' laws. With hydrogen the deviation is scarcely noticeable; with chlorine the deviation is about 1½ per cent.

² Strictly speaking, the density of a gas is the weight of 1 c.c. of the gas at 0° and 760 mm. The density of a gas is usually expressed in terms of a litre of the gas, because the number representing the weight of 1 c.c. would be inconveniently small.

By Avogadro's hypothesis, equal volumes of gases contain the same number of molecules, consequently, the relative density of a gas is proportional to its molecular weight. For, let n represent the number of molecules in a volume v of each of two different gases at the same temperature and pressure (Avogadro's rule), and if the molecules of each gas are alike so that the molecular masses of the one gas can be represented by m_1 , and of the other gas by m_2 , then the mass of the one gas will be nm_1 , and of the other, nm_2 . Let the densities of the two gases be respectively D_1 and D_2 , then, since density denotes the mass of unit volume, $D_1 : D_2 = nm_1/v : nm_2/v$; that is $D_1 : D_2 = m_1 : m_2$.

If we accept this deduction, it enables us to determine the molecular weights of gases, once we have fixed an arbitrary standard for the density. Cannizzaro's unit: hydrogen = 2, is frequently taken as the standard, or as Cannizzaro expressed it, "the quantity of hydrogen contained in a molecule of hydrogen chloride" is taken as unity. The determination of the molecular weight of a gas is thus reduced to a laboratory measurement—the determination of the relative density of the gas. Methods for measuring vapour densities are described in a later chapter. The numerical values for the molecular weight and the relative density of a gas referred to hydrogen (= 2) are the same.¹ That is,

$$\text{Molecular weight} = \text{Relative density (H}_2 = 2) \quad \dots \quad (1)$$

If the density be determined, as is frequently the case, with reference to air = unity, then, since the density of air with reference to $\text{H}_2 = 2$ is 28.75; or with reference to $\text{O}_2 = 32$, 28.98, it follows:

$$\text{Molecular weight} = 28.98 \times \text{Relative density (air} = 1) \quad \dots \quad (2)$$

These very important deductions were made by S. Cannizzaro in a pamphlet published in 1858. Before this step was taken, rank confusion prevailed in chemical literature. The terms "atomic weight," "combining weight," and "molecular weight" were used and abused in every conceivable way. After reading Cannizzaro's pamphlet, Lothar Meyer (1860) said: "the scales fell from my eyes, my doubts disappeared, and a feeling of tranquil security took their place."

Cannizzaro gave the following numbers, among others, for the densities of the different gases referred to hydrogen taken as 2, or to a semi-molecule of hydrogen taken as unity:

	Relative densities.
Hydrogen	2.0
Ordinary oxygen	32.0
Chlorine	71.0
Nitrogen	28.0
Water vapour	18.0
Hydrogen chloride	36.5

If, therefore, the molecules of hydrogen, oxygen, nitrogen, and chlorine

¹ The student must be careful to note that if the unit be $\text{H} = 1$ or $\text{O} = 16$, the relative density = $\frac{1}{2}$ molecular weight; or the molecular weight = $2 \times$ relative density. It is unfortunate that these different units are employed, even though all give the same final result. The questions appended to this and some subsequent chapters have been compiled from different sources, and the different units have not been reduced to one uniform system.

contain two atoms, the atomic weights of these gases will be half the respective molecular weights. Hence:

TABLE II.—CANNIZZARO'S TABLE OF ATOMIC WEIGHTS.

Element.	Relative density of gas.	Atomic Weight Density \div 2.
Hydrogen	2	1.0
Oxygen	32	16.0
Chlorine	71	35.5
Nitrogen	28	14.0

In the case of compounds, if the molecule of hydrogen chloride contains an atom of chlorine and an atom of hydrogen, the molecular weight will be $35.5 + 1 = 36.5$; and the molecule of water vapour containing two atoms of hydrogen and one atom of oxygen, will have a molecular weight of $16 + 2 = 18$. Hence, given the molecular weight of a compound gas, and the weights of the atoms of all but one of the elements, it is possible to compute the weight of the atom or atoms of that element in the molecule in question. The *modus operandi* will be discussed in the next two sections.

Avogadro explicitly guarded against the assumption that the number of constituent atoms must always be 2. There is really nothing in the facts to justify the assumption that the atoms are simple particles. For all we know to the contrary, the atoms may be clusters of n particles. Indeed, we shall soon review some cogent evidence which has led to the inference that Dalton's atoms are not nature's irreducible minima. Even if this inference be valid, each cluster of n particles has a definite weight—atomic weight—and enters into and is expelled from chemical combination as if it were a simple particle. If an atom be a cluster of particles, each cluster, so far as we can tell, has up to the present time behaved in chemical reactions as if it were an individual particle.

Problem.—To deduce Avogadro's law from the relation between the relative densities and the molecular weights of the gases. Let M_1 and M_2 denote the weights of the molecules of two gases—A and B respectively; further, let n_1 and n_2 respectively denote the number of molecules in unit volumes of the two gases. The weights of unit volumes (i.e. the densities) of the two gases will be $M_1 n_1$ and $M_2 n_2$. The observed fact is that the molecular weights (M_1 and M_2) of the gases are proportional to the densities ($M_1 n_1$ and $M_2 n_2$) of the gases; or $M_1 n_1 : M_2 n_2 = M_1 : M_2$, from which it follows that in unit volumes of the two gases $n_1 = n_2$. This is the symbolic way of stating Avogadro's law. Hence, it has been claimed that Avogadro's postulate can be deduced from the relation between the molecular weights and the densities of two gases.

This is a convenient example for cautioning the student not to be misled by the apparent precision and rigorous accuracy conveyed to his mind by reasoning expressed in mathematical symbols. Some affirm, on the strength of the simple demonstration just indicated, that "Avogadro's hypothesis is true." The reasoning is perfectly sound, but what about the premises, or statements upon which the reasoning is based? If the

student has followed the description of Avogadro's work, he will see that the method for the determination of molecular weights tacitly assumes Avogadro's hypothesis is true. Hence, if the mathematical demonstration be employed to *prove* that "Avogadro's hypothesis is true," we argue in a vicious circle. We have assumed in the premises what we sought to "prove" in the demonstration.

§ 4. The Formulæ of Compounds.

Avogadro's hypothesis affords a bridge by which we can pass from large volumes of gases, which we can handle, to the minuter molecules, which individually are invisible and intangible.—W. A. SHENSTONE.

Since Cannizzaro's time, an enormous number of molecular weights have been determined by the vapour density method. If the molecule cannot be decomposed, we must assume that it is composed of one kind of matter only. If the substance is compound, it must be analysed so as to find the ratio, by weight, of its component elements referred to the oxygen standard (16). For instance, suppose that the analysis of a gaseous compound furnished:

Nitrogen 82·35 per cent. ; Hydrogen 17·65 per cent.

Using Cannizzaro's atomic weights, p. 77, oxygen = 16, hydrogen = 1, and nitrogen 14, the compound has its nitrogen and hydrogen atoms in the following proportion by weight:

$$\frac{82\cdot35}{14} \text{ nitrogen atoms : } \frac{17\cdot65}{1} \text{ hydrogen atoms.}$$

That is,

$$5\cdot9 \text{ nitrogen atoms : } 17\cdot65 \text{ hydrogen atoms.}$$

By hypothesis we cannot have fractions of atoms. The nearest whole numbers are 3 hydrogen atoms for one nitrogen atom. Since the sum of the atoms in the compound must represent the molecular weight, it follows that the molecular weight must be $3n + 14n = 17n$. Or the molecular weight is 17×1 ; 17×2 ; 17×3 ; ... or $17n$. The formula is N_nH_{3n} . We can get no further until we know the molecular weight. If the vapour density of the compound (hydrogen = 2) be 17, the molecular weight is 17. Hence, $17 = 17n$, or $n = 1$. The compound analysed can therefore be represented by the formula NH_3 .

EXAMPLES.—(1) As indicated on p. 63, E. W. Morley (1895) found, in some careful experiments on the synthesis of water, that: hydrogen used 3·7198 grms.; oxygen used 29·5335 grms.; water formed 33·2530 grms. That is, one part by weight of hydrogen combines with 7·94 parts by weight of oxygen to produce 8·94 parts by weight of steam. A molecule of steam must contain n atoms of hydrogen, because parts of an atom do not take part in chemical changes. Hence, n parts by weight of hydrogen per 7·94 parts by weight of oxygen give a molecule of steam of weight 8·94 n . This all follows from the atomic theory. To apply Avogadro's hypothesis, with Cannizzaro's standard, the density of the steam must be determined. It lies between 16 and 20. It is difficult to determine the number exactly. If $n = 1$, the density of the steam molecule will be near 8·94. This does not agree with the observed density 16 to 20. If $n = 2$, the density of the steam will be 17·88; and if $n = 3$, the density of steam will be 26·82. Hence, $n = 2$. This means that each molecule of water vapour contains 2 atoms of hydrogen, atomic

weight 1, and one atom of oxygen atomic weight 15.88; or if we make our unit oxygen = 16, the atomic weight of hydrogen will be 1.008.

(2) Two different compounds have the same ultimate composition, namely: carbon 92.31 per cent., hydrogen 7.69 per cent., but the one has a relative density 26, and the other a relative density 78 ($H = 2$). What is the formula of each compound? There are $92.31 \div 12 = 7.7$ carbon atoms per $7.7 \div 1 = 7.7$ hydrogen atoms. But we cannot have fractions of atoms; hence, dividing by 7.7, we get the ratio 1:1. That is, the formula of the compound is C_nH_n . The molecular weights of this series of compounds is $(12 + 1)n$ or $13n$. If $n = 2$, the molecular weight will be 26. Hence, one of the compounds is C_2H_2 , and the other is C_6H_6 .

In calculating formulæ for substances which cannot be vaporized, and one of the methods to be described later cannot be applied, it is usual to assume that the molecule has the simplest possible formula. In that case the formula is said to be *empirical*. Some prefer to use the term "formula weight" in place of "molecular weight" when the actual molecular weight has not been determined. The formula weight, like the molecular weight of a compound, is the sum of the atomic weights of the elements represented by the known or assumed formula of the compound.

EXAMPLES.—(1) 10 grams of pure tin when oxidized in air gave 12.7 grams of oxide. What is the formula of tin oxide? The atomic weight of tin is 119, and of oxygen 16. Hence, the ratio: Tin: oxygen = $10 \div 119 : 2.7 \div 16 = 0.084 : 0.17 = 1 : 2$. The formula is therefore written SnO_2 , although there is nothing to show why it is not Sn_2O_4 ; Sn_3O_6 ; ... Sn_nO_{2n} .

(2) A sample of crystallized sodium carbonate furnished on analysis 37.2 per cent. of Na_2CO_3 , and 62.8 per cent. of H_2O . What is the formula of the compound? The ratio $Na_2CO_3 : H_2O = 37.2 \div 106 : 62.8 \div 18 = 0.35 : 3.49 = 1 : 10$. Hence, the formula is taken as $Na_2CO_3 \cdot 10H_2O$, although there is nothing to show why it is not some multiple of this, say, $nNa_2CO_3 \cdot 10nH_2O$.

§ 5. The Relative Weights of the Atoms.

The atoms are so inconceivably little that their aggregates are alone the ostensible subjects of experiment.—S. BROWN.

It has already been stated that the conceptions "molecular weight" and "atomic weight" are quite independent of our theories about the nature of atoms and molecules; nor are the conceptions much affected by the actual weights of the atoms and molecules, because the terms under consideration are definite expressions of Avogadro's hypothesis coupled with observed facts. It might therefore have been misleading to head this paragraph: "Weighing the Atoms." There are reasons for supposing that the molecular weight of some compounds in the liquid or solid condition is a multiple of the molecular weight of the same substance in the gaseous condition. The molecule of steam approximately corresponds with the formula H_2O ; but in liquid water there are reasons for supposing the molecule is either $(H_2O)_3$ or $(H_2O)_4$, that is, the formula for liquid water is not H_2O , but either H_6O_3 or H_8O_4 .

Refer back to the difficulty in fixing the atomic weight of carbon from the ratio of the weights of carbon and oxygen in the two oxides of carbon which we encountered in applying Dalton's atomic theory. Suppose that we do not know the atomic weight of carbon, but that we do know the composition of a number of volatile carbon compounds as well as their relative densities or molecular weights, Table III.

TABLE III.—MOLECULAR WEIGHTS OF SOME CARBON COMPOUNDS.

Volatile compound of carbon.	Composition by weight.	Molecular weight.	Amount of carbon per molecule.
Carbon monoxide . .	Carbon 12; oxygen 16	28	12
Carbon dioxide . .	Carbon 12; oxygen 32	44	12
Methane	Carbon 12; hydrogen 4	16	12
Ethylene	Carbon 24; hydrogen 4	28	$12 \times 2 = 24$
Propylene	Carbon 36; hydrogen 6	42	$12 \times 3 = 36$
Carbon disulphide . .	Carbon 12; sulphur 64	76	12

The smallest weight of carbon in a molecule of any of its known compounds is 12, and consequently this number is assumed to be the atomic weight of carbon. The atomic weights of a great number of the elements have been determined in a similar manner.

The actual method used in finding the atomic weight of an element thus involves:

- (1) An exact analysis of a compound containing the given element; and consequently the compound investigated must be one which lends itself to exact analysis.
- (2) The compound must be one which can be prepared in a highly purified condition.
- (3) The compound must be volatile without decomposition, so that its vapour density can be determined.
- (4) The compound must contain the smallest proportion of the element under investigation. This matter may need further amplification.

In 1894, J. A. Wanklyn claimed to have discovered a series of hydrocarbons, one member of which contained carbon 102 parts by weight, and hydrogen 17 parts, and had a vapour density of nearly 116 (hydrogen 2). Assuming the atomic weight of carbon is 12, and of hydrogen 1, these numbers give formula $C_{3.5}H_{17}$. If this statement had been corroborated, and we were quite sure that Wanklyn's hydrocarbons were not mixtures, it would be necessary to make the atomic weight of carbon = 6, and write the formula of the compound in question $C_{17}H_{17}$, and this in spite of the fact that thousands of compounds of carbon are known, and all agree with the number 12 for the atomic weight of carbon. The formula of carbon monoxide—CO—would then be written C_2O , etc. But Wanklyn's claim has never been established.

These remarks emphasize the importance of examining as large a number of volatile compounds as possible when fixing the atomic weight of an element. If only a small number of compounds be examined, there is always a possibility, and perhaps a probability, that the actual minimum weight does not occur amongst the set of compounds taken. It follows, therefore, that the atomic weight of an element is the least amount of that element present in any molecule of all its known volatile compounds. The value so obtained is the maximum possible value; the real value may afterwards prove to be a submultiple of this. The atomic weight must be a whole multiple or submultiple of its combining weight. Owing to the fact that the molecular weights of so many

volatile compounds of carbon are known, it is not very probable that the atomic weight of carbon is less than 12.

There are several other methods of computing molecular and atomic weights of the different elements. Fortunately, atoms and molecules possess other qualities besides mass which are dependent upon their "atomic weights" and which can be readily measured. Some of these will be described later.

§ 6. The Elements.

Every chemical element is regarded as having a distinct nature of its own, which nature, moreover, determines all its activities.—B. P. BROWNE.

What are the best representative values for the atomic weights of the elements?—The best available determinations of the value of the oxygen-hydrogen ratio give numbers ranging between 1.005 and 1.008 when the standard of reference is oxygen 16. All measurements made by man are affected by unavoidable errors of experiment; and measurements of the numerical value of all constants differ within certain limits amongst themselves (see p. 28). It is convenient to select one representative value from the set of different observations ranging between the limits 1.005 and 1.008. The majority of chemists have agreed to let the International Committee of Atomic Weights decide what are the best representative values for the atomic weights of all the elements year by year. Hence, the generally accepted ratio for the atomic weights of hydrogen and oxygen is 1.008 : 16. Every time new and more refined methods of measurement are employed, a change—generally insignificantly small—may be necessary. The student must recognize that the true atomic weights cannot be altered by the votes of the majority of the members on the International Committee of Atomic Weights. There is an uncertain factor in the accepted values of the atomic weights, as there is in all our judgments. Aristotle was no doubt right: "Nothing can be positively known, and even this cannot be positively asserted." This doctrine, however, if rigorously applied, would paralyse all action. Accordingly, sound-minded people are accustomed to balance the evidence and then act. A careful consideration of all the available evidence considerably reduces the risk of error, and this method, adopted by the Committee, appears to be the most satisfactory solution of the problem.

The atomic weights of a few of the more important elements are indicated in the table on next page. The numbers are those recommended by the International Committee on Atomic Weights. The full table appears inside the front cover of this book.

For ordinary calculations involving the use of atomic weights, all the atomic weights, excepting chlorine (35.5), copper (63.5), nickel (58.5), and zinc (65.5), are rounded off to the nearest whole numbers. The elements just named are then assigned the atomic weights indicated in the brackets. Some chemists—G. D. Hinrichs, for example—firmly believe that the rounded numbers are the best representative values of the atomic weights, and that the small deviations from the rounded numbers indicated in the "International Table" represent real, if unrecognized, errors of experiment. In view of the work on isotopes, to be discussed in the concluding chapters, this question has lost its significance until we have learned to

TABLE IV.—INTERNATIONAL ATOMIC WEIGHTS. O = 16.

Aluminium	Al	27.1	Iron	Fe	55.84
Antimony	Sb	120.2	Lead	Pb	207.20
Arsenic	As	74.96	Magnesium	Mg	24.32
Barium	Ba	137.37	Manganese	Mn	54.93
Bismuth	Bi	208.0	Mercury	Hg	200.6
Boron	B	11.0	Nickel	Ni	58.68
Bromine	Br	79.92	Nitrogen	N	14.01
Calcium	Ca	40.07	Oxygen	O	16.00
Carbon	C	12.005	Phosphorus	P	31.04
Chlorine	Cl	35.46	Platinum	Pt	195.2
Chromium	Cr	52.0	Potassium	K	39.10
Cobalt	Co	58.97	Silicon	Si	28.3
Copper	Cu	63.57	Silver	Ag	107.88
Fluorine	F	19.0	Sodium	Na	23.00
Gold	Au	197.2	Sulphur	S	32.06
Hydrogen	H	1.008	Tin	Sn	118.7
Iodine	I	126.92	Zinc	Zn	65.37

interpret the fact that many of the atomic weights at present accepted are averages of mixtures of two or more different kinds of atoms which at present are given the same name, although they have slightly different atomic weights. So far as has yet been observed, these atoms cannot be distinguished from one another by any chemical process yet tried.

Why is oxygen = 16 taken as the standard?—During the latter part of the 19th century, J. Dalton's (1803) standard, hydrogen = 1, was used for the atomic weights instead of oxygen = 16. Hydrogen was selected because it is the lightest element known. J. S. Stas (1860-65) pointed out that the determination of the atomic weight of an element should be connected with the standard as directly as possible. Very few compounds of the metals with hydrogen are suitable for an atomic weight determination, while nearly all the elements form stable compounds with oxygen. Hence, if hydrogen be the standard, it is necessary to find the exact relation between the given element and oxygen, and then calculate what that relation would be on the assumption that the relation between hydrogen and oxygen is known. Every improved determination of the relation between hydrogen and oxygen would then be followed by an alteration in the atomic weight of every other element whose value, with respect to hydrogen as a standard, has been determined by the indirect process just indicated. The determination of the exact relation between hydrogen and oxygen appears to be more difficult than many other determinations, and hence, the majority of chemists think it better to refer the atomic weights of the elements to oxygen = 16 as the standard instead of making the atomic weights depend on the more or less uncertain relation H:O. The standard oxygen = 16 is quite arbitrary. T. Thomson (1825) used oxygen = 1; W. H. Wollaston (1814), 10; J. S. Stas (1860-65), 16; and J. J. Berzelius (1830) used oxygen = 100 as standard. The latter number makes the atomic weights of many elements inconveniently large, and if the atomic weight of oxygen be any whole number less than 16, fractional atomic weights will be required. The use of the "oxygen-16" unit involves the least change in the numbers which were in vogue when "hydrogen-unity" was the standard.

Division of the elements into metals and non-metals.—It is often convenient to divide the elements into two groups: metals and non-metals. Like most systems of classification an exact subdivision is not possible, because some elements exhibit properties characteristic of both classes. Very roughly, the properties of the metals can be contrasted against the properties of the non-metals as indicated in the subjoined scheme:

TABLE V.—THE PROPERTIES OF THE METALS AND NON-METALS CONTRASTED.

Metals.	Non-metals.
1. Form basic oxides.	1. Form acidic oxides.
2. Generally dissolve in mineral acids, giving off hydrogen.	2. Do not usually dissolve easily in mineral acids.
3. Either form no compounds with hydrogen, or form unstable compounds—usually non-volatile.	3. Form stable compounds with hydrogen—these are usually volatile.
4. Solid at ordinary temperature (excepting mercury).	4. Gases, liquids or solids at ordinary temperatures.
5. Usually volatilize only at high temperatures.	5. Excepting carbon, boron, and silicon, the non-metals are either gaseous or volatilize at low temperatures.
6. When in bulk the metals reflect light from polished or freshly cut surfaces.	6. Do not usually reflect light very well.
7. Specific gravity is generally high.	7. Specific gravity generally low.
8. Good conductors of heat and electricity. Electrical resistance usually increases with rise of temperature.	8. Bad conductors of heat and electricity. Electrical resistance usually decreases with rise of temperature.
9. More or less malleable and ductile.	9. Malleability and ductility are not well defined.
10. Molecules usually monatomic in vaporous state.	10. Molecules usually polyatomic in vaporous state.

To show how difficult it is to draw a hard-and-fast line of demarcation between metals and non-metals, the non-metals arsenic, antimony, and tellurium would be classed with the metals if we depended exclusively upon 6, 7, and 8; hence, some introduce a third division—the metalloids—to include the hybrids, or elements which have properties characteristic of both the metals and the non-metals. The metals lithium, sodium, potassium, magnesium, and aluminium have a low specific gravity. The non-metals carbon, boron, and silicon are less volatile than most metals. The non-metal hydrogen is a good conductor of heat; and the non-metal graphitic carbon is a good conductor of heat and electricity. Hence the division of the elements into metal and non-metals is but a rough system of classification, arbitrarily adopted because it is convenient.

In all systems of classification, the attempt is made to bring together in one group the things which are alike in general properties, and to separate those which are unlike. The attempt to group the elements by a code of definitions is foredoomed to failure. There is a seductive simplicity about a definition which may be attractive, but it is artificial and often misleading. As T. Campanella (1590) expressed it: "Definition is the

end and epilogue of science. It is not the beginning of our knowing, but only of our teaching."

§ 7. The Relation between the Molecular Weights and the Volumes of Gases.

The theory of molecules is an ideal conception placed by the mind like another Atlas underneath a measureless world of facts to give them intelligible cohesion and hold them up to view.—S. BROWN.

The molecular weight of any gas is numerically equal to the weight of any volume of the gas when the weight of an equal volume of hydrogen under the same physical conditions of temperature and pressure is 2. Two grams of hydrogen, taken as the standard, occupy 22.3 litres at normal temperature—0°—and normal pressure—760 mm. of mercury. Hence, it follows directly from Avogadro's hypothesis that the molecular weight of any gas, expressed in grams, occupies, approximately, 22.3 litres at 0° and 760 mm. pressure. Consequently, to find the molecular weight of a gaseous substance, weigh 22.3 litres of the gas at a convenient temperature and pressure; calculate the corresponding volume at 0° and 760 mm. pressure, and calculate by proportion the weight of 22.3 litres.

EXAMPLE.—A litre of gas at 20° and 730 mm. weighs 1.764 grams: what is the molecular weight of the gas? By the method of calculation indicated in the next chapter, one litre of a gas at 20° and 730 mm. pressure contracts to 894.5 c.c. at 760 mm. and 0°. Hence, if 894.5 c.c. weigh 1.764 grams, 22.3 litres will weigh 43.97 grams. Hence the molecular weight of the gas is nearly 44.

It must here be mentioned that the number 22.3 is not quite right for all gases. Many gaseous molecules have a slight attraction for one another, so that the molecules are slightly more closely packed than is represented by Avogadro's hypothesis. The greater the intermolecular attraction, the greater the weight of 22.3 litres, and consequently, the less the volume of a molecular weight of the gas expressed in grams. Thus, experiment shows:

Hydrogen	Oxygen	Nitrogen	Chlorine	Hydrogen chloride	Carbon dioxide	Steam (0°, 760 mm.)	Mercury
22.40	22.39	22.45	22.01	22.22	22.26	22.39	22.55

The deviation from 22.3 can be neglected in ordinary chemical calculations.

The molecular weight of a compound not only tells us a weight, but it also tells us that if the molecular weight be expressed in grams the gas will occupy 22.3 litres at 0° and 760 mm. Further, the molecular weight of a gas, expressed in kilograms, occupies, approximately, 22.3 cubic metres at 0° and 760 mm. pressure. By mere chance, the number of avoirdupois ounces in a kilogram is 35.26, which is very nearly the same as the number of cubic feet in a cubic metre (35.31)—J. W. Richards. The difference is only one-seventh of one per cent. Hence, the molecular weight of any gas, expressed in avoirdupois ounces, occupies approximately 22.3 cubic feet at 0° and 760 mm. pressure. These factors are useful in calculations involving cubic feet, cubic metres, and litres.

§ 8. Chemical Equations and Chemical Arithmetic.

In his calculations, the chemist relies on the supposed numerical relations of the invisible, intangible, and immeasurable particles he calls atoms. These relations have been determined by others in whom he has confidence, and the accuracy of these constants has to be accepted on faith.—H. C. BOLTON.

All (chemical) change is an effect of the union or separation of atoms.—LUCRETIVS.

When the initial and final products of a chemical reaction as well as the composition and proportions of the molecules concerned in the reaction are known, the facts can usually be symbolized in the form of a chemical equation. There are some limitations which will be described later.

1. The molecular weight of an element or compound is the sum of the atomic weights of all the atoms of the constituent elements.—Thus, if the atomic weight of oxygen is 16, of hydrogen 1, of nitrogen 14, of sulphur 32, and of iron 56, the molecular weight of water, H_2O , is 18; that of sulphuric acid, 98; and that of ferrous ammonium sulphate, $FeSO_4 \cdot (NH_4)_2SO_4 \cdot 6H_2O$, 392—since $56 + 32 + 4 \times 16 + 2(14 + 4) + 32 + 4 \times 16 + 6(2 + 16) = 392$.

2. The equation indicates the nature of the atoms and the supposed composition of the molecules concerned in the reaction; as well as the proportions of the different molecules in the initial and final products of the reaction.—For instance, when mercury is heated in air, and mercuric oxide, HgO , is formed, the reaction can be represented in symbols: $2Hg + O_2 = 2HgO$. We here ignore the nitrogen of the air because, so far as we can tell, it plays no direct part in the chemical reaction. Similarly, when mercuric oxide is heated to a high temperature, it decomposes, forming metallic mercury and oxygen. In symbols, $2HgO = 2Hg + O_2$. The symbol “=” is used instead of the words “produces” or “forms,” and the symbol “+” is used for “together with” on the right side of the “=” sign, and for “reacts with” on the left side. The latter equation reads: “Two molecules of mercuric oxide, on decomposing, produce a molecule of oxygen and two molecules of monatomic mercury.” The number and kind of the atoms of the two sides of the equation must always be the same (persistence of weight).

3. The equation indicates the proportions by weight of the substances concerned in the reaction.—The atomic weight of mercury is 200, and the atomic weight of oxygen is 16, hence, the molecular weight of mercuric oxide is 216, and of oxygen 32. The latter equation can therefore be read: “432 grams (ozs. or tons) of mercuric oxide, on decomposing, form 32 grams (ozs. or tons) of oxygen gas and 400 grams (ozs. or tons) of metallic mercury. Hence, the chemical equation can be employed in all kinds of arithmetical problems dealing with weights of substances formed or produced.

EXAMPLES.—(1) How much mercuric oxide is required to furnish 20 grams of oxygen gas? Write down the proper equation; write 432 below the mercuric oxide, and 32 below the oxygen. We are not concerned with the mercury in this problem. Since we read, from the equation: 32 grams of oxygen are furnished by 432 grams of mercuric oxide, 1 gram of oxygen will be furnished by $432 \div 32 = 13.5$ grams of mercuric oxide; and 20 grams of oxygen will come from $20 \times 13.5 = 270$ grams of mercuric oxide.

(2) How much oxygen can be obtained from 30 grams of mercuric oxide? Answer: $2\frac{1}{2}$ grams.

4. The equation indicates the proportions by volume of the gases concerned in the reaction.—We have seen in the preceding section that if we express

Molecular weights in	Volume at 0° and 760 mm. per molecular weight.
Grams	22.3 litres.
Kilograms	22.3 cubic metres.
Ozs. (avoir.)	22.3 cubic feet.

Consequently, we can express the idea conveyed by the equation, $2\text{HgO} = \text{O}_2 + 2\text{Hg}$, in these words: "432 grams (kilograms or ozs.) of mercuric oxide will furnish 32 grams (kilograms or ozs.) of oxygen, or 22.3 litres (cub. metres or cub. ft.) of oxygen gas at 0° and 760 mm., and 400 grams (kilograms or ozs.) of mercury."

EXAMPLES.—(1) What volume of oxygen will be obtained by heating 30 grams of mercuric oxide? 432 grams of mercuric oxide will furnish 22.3 litres of oxygen gas, hence 30 grams will furnish $30 \times 22.3 \div 432 = 1.55$ litres of oxygen gas at 0° and 760 mm. pressure.

(2) How much mercuric oxide will be needed for 10 cub. ft. of oxygen gas at 0° and 760 mm. pressure? Here 22.3 cub. ft. of the gas come from 432 ozs. of mercuric oxide, hence $432 \times 10 \div 22.3 = 193$ ozs., or 12 lbs. 1 oz. of mercuric oxide are required.

§ 9. The Relation between Atomic and Combining Weights—Valency.

When the formulæ of inorganic compounds are examined, even a superficial observer is struck by their general symmetry. Without offering any hypothesis as to the cause of this symmetrical grouping of the atoms, it is sufficiently evident that such a tendency exists, and that the combining power of the atoms of the attracting elements is always satisfied by the same number of atoms without reference to the chemical behaviour of the uniting atoms.—E. FRANKLAND.

Observation shows that the relative combining weights of oxygen and hydrogen are very nearly as $\text{O} : \text{H} = 8 : 1$; and that the atomic weights of oxygen and hydrogen, deduced from the atomic theory and Avogadro's hypothesis, are very nearly as $\text{O} : \text{H} = 16 : 1$. In fine, the atomic weight of oxygen is twice its combining weight. For carbon in carbon dioxide we have the combining weight 3, while the atomic weight of carbon is 12, that is, the atomic weight of carbon is four times the combining weight. In the case of hydrogen and chlorine the atomic and combining weights are the same. The number of times the combining weight or equivalent is contained in the atomic weight is called the valency of the element. Hence,

$$\frac{\text{Atomic weight}}{\text{Combining weight}} = \text{Valency.}$$

This means that when the combining or equivalent weight of an element is multiplied by an integer representing the valency of the element, the product is the atomic weight.

The meaning of valency can be represented another way, for valency also represents a "habit" of an element for combination; it has nothing to do with the force holding the atoms together. The valency of an element is obtained by finding—directly or indirectly—how many atoms of hydrogen can combine with or be replaced by an atom of the given element. The valency of hydrogen is always taken as unity. Hence the

definition: The valency of an element is a number which expresses how many atoms of hydrogen, or of other atoms equivalent to hydrogen, can unite with one atom of the element in question.

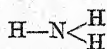
Nomenclature.—With hydrogen and chlorine the atomic and combining weights are the same, and the valency is unity. These elements are accordingly said to be univalent, or monads; for similar reasons oxygen is bivalent, or a dyad; nitrogen is trivalent, or a triad; and carbon is quadrivalent, or a tetrad.

The valency of an element is frequently represented by attaching the necessary number, in dashes or Roman numerals, to the top right-hand corner of the symbol for the element, as suggested by W. Odling in 1855. Thus, the symbols H^I and Cl^I respectively mean that hydrogen and chlorine are univalent; O^{II} means that oxygen is bivalent; N^{III} means that nitrogen is trivalent; and C^{IV} that carbon is quadrivalent. By collecting together a few compounds with their symbols, the idea can be made clearer.

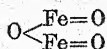
Univalent.	Bivalent.	Trivalent.	Quadrivalent.	Quinquevalent.	Sexivalent.
H^ICl^I	$H_2^IO^{II}$	$H_3^IN^{III}$	$H_4^IC^{IV}$	$P^V F_5^I$	$S^{VI} F_6^I$
Na^ICl^I	$Mo^{II} Cl_2^I$	$Fe^{III} Cl_3^I$	$C^{IV} O_2^{II}$	$W^V Br_5^I$	$U^{VI} F_6^I$
K^IH^I	$Zn^{II} Cl_2^I$	$Mo^{III} Cl_3^I$	$Mo^{IV} Cl_4$	$Mo^V Cl_5^I$	$Mo^{VI} F_6^I$

Some heptads and octads are known. The elements generally combine in such a way that an equal number of valencies are opposed to one another.

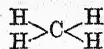
Structural, graphic, or constitutional formulæ.—The valency of an element is sometimes represented by attaching the necessary number of hyphens to the symbol for the element. This enables the molecules of a substance to be represented by a kind of graphic formula. The symbol for hydrogen will have one hyphen; oxygen, two; nitrogen, three; carbon, four; etc. The symbol for hydrogen chloride then becomes $H-Cl$; potassium iodide, $K-I$; water, $H-O-H$; mercuric oxide, $Hg=O$; a molecule of hydrogen, $H-H$; a molecule of oxygen, $O=O$; carbon dioxide, $O=C=O$; and



Ammonia.



Ferric oxide.



Methane.

Accordingly, the term "bonds" or "links" is sometimes employed instead of "valencies."¹

Graphic formulæ are also called structural or constitutional formulæ. *Structural formulæ primarily assume that the chemical properties of a substance are determined by the arrangement of the atoms in the molecules; and if the molecules of two compounds of the same chemical composition have their atoms differently arranged, the properties of the two compounds will be different.* Graphic formulæ are sometimes very convenient for representing the composition of compounds, but the student would err rather seriously

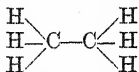
¹ The hyphens are generally attached so that the graphic formula occupies as little space as possible. I believe A. S. Couper first used linking bars between the symbols of the combining elements in 1858; and the present system developed in its present form through the work of A. Kekulé (1859), C. Brown (1865), and E. Frankland (1866).

if he supposes that the symbol given above for, say, methane represents the way the atoms are actually grouped in the molecule of methane. This would involve a leap far beyond our real knowledge. In some cases, however, the little knowledge we do possess can be better summarized by a graphic formula than in any other way, and the graphic formula furnishes a clearer mental image of the curious way certain groups of atoms remain clustered together through a complex series of chemical changes than if the reaction were represented by ordinary symbols. "Hence," said J. F. Heyes, "the language of valency is useful in focussing the mysteries enshrined within the chemical molecule." A graphic formula is thus a kind of "dummy" model to illustrate the way a compound is formed, how it decomposes, and the relations between one compound and another. The student must not believe for one moment that the model simulates reality. The remarkable work which has been done by the aid of structural formulae will always justify their use in the past and present, whatever future generations may think of them.¹ There are, however, many compounds which do not fit the valency hypothesis very well. When experiments are guided by a wrong theory, only those things sought are likely to be found. The hypothesis may not have any counterpart in nature, and, as A. G. V. Harcourt has said, if a chemist were to depart from the general course, and attempt to prepare compounds not indicated by theory, he would perhaps obtain results of more than usual interest.

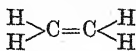
Maximum and active valency.—Most elements have more than one valency. Stannous oxide has a composition corresponding with SnO , and stannic oxide, SnO_2 . In the former case the tin is said to be bivalent; and in the latter, quadrivalent. There are thus two series of tin compounds—stannous and stannic. Similarly with copper, iron, etc. There are also two carbon oxides, carbon monoxide, CO , and carbon dioxide, CO_2 . If carbon monoxide could be written $\text{O}=\text{C}=\text{C}=\text{O}$, and there is nothing in the analysis by weight which prevents this, all might be well; but writing the formula in this manner would involve a contradiction of Avogadro's hypothesis, since the vapour density of carbon monoxide corresponds with the molecule CO , not C_2O_2 . We cannot see the way clear to admit carbon monoxide as an exception to Avogadro's hypothesis, for that would introduce confusion into our system, and there would be no immediate prospect of restoring order. Some get over the difficulty by assuming that two of the free valencies in carbon monoxide mutually saturate one another, and write the graphic formula $\text{O}=\text{C} \angle$; others assume that oxygen is quadrivalent, and write $\text{C} \equiv \text{O}$. The case

¹ The student will find valency to be a useful aid in remembering the composition of compounds of different elements. Given the valency of 12 elements, each of which can form one compound with 12 other elements of known valency, it is possible to write down the formulae of 144 compounds which would be very probably in harmony with the known laws of chemical combination. It would be said that MgCl is not a probable compound of magnesium and chlorine, nor is MgCl_3 ; the correct way of writing the combination is MgCl_2 . The student should therefore remember the valency of each element he studies; and in that way, much that appears confused and disorderly will seem methodical and regular. The valency of an element may be recalled by reference to its compounds with other elements of known valency. Thus, the chemist does not memorize valencies themselves, but recovers them when needed by recalling a familiar compound or compounds. If the valency of magnesium were forgotten and magnesium chloride, MgCl_2 , were remembered, then, if chlorine be univalent, magnesium is bivalent,

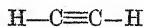
of sulphur bivalent in hydrogen sulphide, H-S-H ; quadrivalent in sulphur dioxide O=S=O ; and sexivalent in sulphur trioxide: $\text{O} \begin{smallmatrix} \diagup \\ \diagdown \end{smallmatrix} \text{S} \begin{smallmatrix} \diagup \\ \diagdown \end{smallmatrix} \text{O}$, fits very well into this scheme. So do the series of compounds represented by ethane, C_2H_6 ; ethylene, C_2H_4 ; and acetylene, C_2H_2 , which are respectively represented by the graphic formulæ:



Ethane.



Ethylene.



Acetylene.

It has been supposed that valency is a "fundamental property of the atom which is just as constant and invariable as the atomic weight"—A. Kekulé; and further that each element has a maximum valency towards certain other elements. When an element in a compound appears to have a lower valency than its maximum valency, the compound is said to be an unsaturated compound, in contrast with a saturated compound in which the atoms are exercising their maximum valency. In many unsaturated compounds the valencies appear to diminish in pairs. The pairs of "sleeping valencies" or "latent bonds" are supposed to be self-saturated. As a matter of fact, the hypothesis of the self-saturation of the bonds in pairs breaks down completely. The idea probably arose from the application of an inaccurate hypothesis which is stated in some of the older books on chemistry in words like these: "All chemical evidence shows that a body with unsatisfied bonds cannot exist by itself." All chemical evidence, as we shall see, shows nothing of the kind. Mercury and many other elements, when vaporized, give gases with one-atom molecules.

The principle of self-saturation breaks down when applied to the nitrogen oxides, say $\text{N}^{\text{III}}\text{O}^{\text{I}}$. The relative density of the gas (Avogadro's hypothesis) will not let us write N_2O_2 , that is, $\text{O}=\text{N}-\text{N}=\text{O}$. We are therefore confronted with what appears to be an odd unsaturated valency in the molecule $-\text{N}=\text{O}$. Again, molybdenum forms a series of compounds with univalent chlorine or fluorine $-\text{MoCl}_2$, MoCl_3 , MoCl_4 , MoCl_5 , and MoF_6 ; and vanadium forms VCl_2 , VCl_3 , VCl_4 , and VCl_5 . In view of facts like these, it is difficult to maintain the thesis that the apparent inconstancy of the valency of an element is due to the mutual "saturation" of pairs of valencies. Either a molecule can exist with free valencies, or Kekulé's maximum valency hypothesis breaks down when confronted with facts.

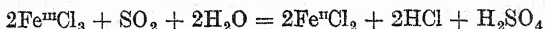
A great many ingenious hypotheses, more or less satisfactory, have been suggested to explain the difficulties. At present we are compelled to frankly admit with W. Lossen (1880) and A. Claus (1881) that the active valency of an element is a variable habit of combination. An explanation of the meaning of valency is thus left open. To distinguish between the greatest valency an element is known to exhibit, and the valency which actually prevails in a particular compound, the terms maximum valency and active valency may be respectively employed. So far as we can see, the active valency of an element is dependent upon the properties of the atoms of the other elements with which it is combined, as well as on the prevailing physical and chemical conditions

to which the element is exposed. Indeed, active valency has been compared with friction in so far as it appears to be called into play by external causes which may vary from zero upwards.

Effect of external conditions on the valency of an element.—We have just stated that the valency of an element is determined by the physical and chemical conditions under which the element is placed. For instance, valency generally diminishes with rise of temperature, *e.g.*, sulphur trioxide, SO_3 , when heated dissociates into sulphur dioxide, SO_2 , and oxygen; and carbon dioxide, CO_2 , into carbon monoxide, CO , and oxygen. Changes in the valency of an element are often induced by oxidizing or reducing agents. Thus, ferrous chloride, FeCl_2 , is oxidized to ferric chloride, FeCl_3 , by the action of hypochlorous acid, HClO :



and, in aqueous solution, ferric chloride is reduced to ferrous chloride by the action of sulphur dioxide



At the same time, it will be noticed, the sulphur dioxide is oxidized to sulphur trioxide $\text{S}^{\text{IV}}\text{O}_2 + \text{O} = \text{S}^{\text{VI}}\text{O}_3$. Hence, oxidation usually involves an increase in the valency of an element, and reduction a decrease. This subject is discussed later.

History.—In the early days of the atomic theory, atoms were all supposed to have an equal capacity for combination with one another. With the growth of the idea of multiple proportions, and the custom of referring the constitution of compounds to certain types— HCl , H_2O , H_3N , H_4C —supposed to be fundamental, the idea of valency gradually became clear, and it was specially emphasized by E. Frankland in 1851.

Various terms were used for "valency" during the clarification of the concept—*e.g.*, "saturation capacity," "combining capacity," "affinity units," "affinity of degree," "basicity," "atomicity," etc.¹ A. W. Hofmann employed the term "quantivalency" in 1865; and this was shortened to "valency" by H. Wichelhaus, in 1868. The term "valency" (or "valence") is now in general use.

§ 10. Radicals or Radicles.

For the chemist, each molecular compound is proximately made up of less compound atoms which are indivisible by forces which can divide their product, and these in turn can be separated by chemical agents into simple atoms.—S. BROWN.

In 1815 J. L. Gay-Lussac, after studying the properties of hydrocyanic acid, reported cyanogen, CN , to be "a remarkable example, and at present a unique example, of a body which, although a compound, plays the part of a single body in its combinations with hydrogen and the metals." Since then a great number of similar groups have been found. For convenience, they are commonly called "radicals," or, following the custom of the London Chemical Society, "radicles." The word "radical"

¹ The term "atomicity" is best reserved to express the number of atoms in the molecule of an element; and "basicity" for the number of stages in which the replaceable hydrogen of an acid can be substituted by a metal.

was previously employed by G. de Morveau and by A. L. Lavoisier with a different meaning. The definition: a radicle is a group of atoms which can enter into and be expelled from combination without itself undergoing decomposition, is virtually that given by J. von Liebig in 1838. Each radicle has its own valency; each acts as an unchanging constant in a series of compounds; and each can be replaced by an element or elements, of like or equivalent valency. A few examples of radicles of different valency may be quoted: Monad radicles—OH, CN (generally written "Cy"), NO_3 , NH_4 (sometimes written "Am"), CH_3 , etc. Dyad radicles— SO_4 , SO_3 , CO_3 , SiO_3 , etc. Triad radicles— PO_4 , $\text{Fe}^{\text{m}}\text{Cy}_3$, etc. Tetrad radicles— $\text{Fe}^{\text{c}}\text{Cy}_6$, SiO_4 , etc. In very few cases has it been possible to isolate the radicle, but the definition has nothing to say about the independent existence of radicles. "Radicles," said A. Kekulé (1858), "are not firmly closed atomic groups, but they are merely aggregates of atoms placed near together which do not separate in certain reactions, but fall apart in other reactions." For convenience, the term "radicle" is sometimes applied to an atom in a compound which can be replaced by another atom or radicle without a further change in the nature of the compound; in that case, the radicle is said to be a "simple radicle" in contrast with "compound radicles," which are "groups of atoms."

Questions.

For drilling students in the arithmetic of chemistry, it is best to use a special book, e.g. S. Lupton, "Chemical Arithmetic," London, 1892; R. L. Whiteley, "Chemical Calculations," London, 1892; J. Waddell, "The Arithmetic of Chemistry," New York, 1899; C. Baskerville and W. L. Eastabrooke, "Progressive Problems in General Chemistry," New York, 1910; C. J. Woodward, "Arithmetical Chemistry," London, 1895; H. W. Bausor, "Chemical Calculations," London, 1914; etc. A number of arithmetical problems are scattered among the questions appended to subsequent chapters of this book.

1. Describe experiments which illustrate the law of combination of gases by volume. It was at one time thought that this law was accounted for by assuming that "equal volumes of all gases at the same temperature and pressure contain the same number of atoms." What facts show that this assumption is incorrect? What hypothesis is now accepted?—*Univ. North Wales*.

2. State some of the facts relating to the union of gases by volume. State the law of combining volumes. Who discovered the law? What hypothesis was propounded to account for the facts underlying the law? What important conclusion follows from this hypothesis?—*Princeton Univ., U.S.A.*

3. It is usually stated that the valency of an element, when variable, differs by two units. Thus the valency of carbon is expressed by the numbers 2 and 4; and of phosphorus 3 and 5. Mention any exceptions to this law and discuss their bearing on the hypothesis that the valency of a body is a fixed and definite quantity.—*London Univ.*

4. A tube contains 45 c.c. of hydrogen and 20.25 c.c. of oxygen, at a temperature of 120°C . In what respects do the contents differ from steam at the same temperature? What effects would be observed on passing a spark through the mixture.—*Cambridge Senior Locals*.

5. Formerly the atomic weight of oxygen was reckoned as 100. With this standard calculate the atomic weight of hydrogen.—*Coll. of Preceptors*.

6. Discuss the question whether $\text{H} = 1$ or $\text{O} = 16$ should be used as the standard for the atomic weights.—*Board of Educ.*

7. Which of the following gases are lighter, and which are heavier than air O_2 , CO , CH_4 , NH_3 , SO_2 , H_2S ? Calculate the weight of 10 litres of CO_2 at N.T.P.—*Board of Educ.*

8. Show that the symbol H_2O best represents the formula for water quite independent of the atomic theory. Hints. The combining weights of hydrogen and oxygen in water are as 1 : 8; or 2 : 16; etc. The formula weight of a gas

is the weight in grams of 22.3 litres of the gas at standard temperature and pressure. This gives a value approaching 18 as the formula weight of water. Experiment also gives 32 as the formula weight of oxygen, O_2 , when the formula weight of hydrogen $H_2 = 2$ is taken as the standard of reference. The formula weight of a compound is the sum of the symbol weights of the elements in the formula of the compound. The combining volumes of hydrogen and oxygen are as 2:1. The only possible formula for water, consistent with these conditions is H_2O . If the formula were HO , the equivalent formula weight would be 17; if H_2O , 18; if H_4O , 20; if H_2O , 33; etc. Hints—hydrogen may be removed in two stages, the oxygen only in one stage.

9. Name two important respects in which metals differ from non-metals. Name an element which may be regarded both as metal and as non-metal, and give some of the reasons.—*Princeton Univ., U.S.A.*

10. Explain in words the precise meaning of the expression " $Mg + H_2SO_4 = MgSO_4 + H_2$ " ($Mg = 24$, $S = 32$, $O = 16$).—*Punjab Univ.*

11. According to C. R. A. Wright (1874), "neither the molecular theory nor the atomic theory, generally so-called, is taken for granted in the formation of chemical equations." Discuss this quotation.

12. What properties may be regarded as characteristic of metals generally? What do you understand to be the meaning of metalloid?—*London Univ.*

13. What is meant by the "equivalent" of an element, and what relation does it bear to the atomic weight?—*London Univ.*

14. State the general characters of metals as a class, and show in what respects they differ from the non-metals.—*Science and Art Dept.*

15. Explain the reasons for believing that "the molecular weight of any gas is twice its density compared with hydrogen."—*Madras Univ.*

16. If an unknown element were given to you, and you were asked to find out whether it should be classed as a metal or non-metal, how would you proceed?—*London Univ.*

17. An element may exhibit several combining or equivalent weights, but only one atomic weight. Explain this phenomenon clearly and briefly.

CHAPTER V

THE PHYSICAL PROPERTIES OF GASES

§ 1. The Atmosphere.

The generality of men are so accustomed to judge of things by their senses that because air is invisible they ascribe but little to it, and think of it as but one remove from nothing.—R. BOYLE (1673).

THE atmosphere in which we live and breathe is really a part of the globe on which we stand. We are not surrounded by mere empty space. On the contrary, we live and move at the bottom of a vast ocean of air, which is just as material as the water which surrounds the flat-fish living at the bottom of the sea. Air was once considered to be a thin, pellucid, evanescent, inscrutable, and imponderable spirit—the spirit of life. Even to-day, air is still used as a symbol for what is spiritual and divine; but to early man the analogy between the impalpable breath of the physical heavens and the inscrutable spirit of God was very real. It was quite a long time before air was recognized to be a gravic material essentially ponderable like earth and sea.

The physical properties of air were studied long before its chemical properties were investigated. Anaxagoras, who lived about the sixth century B.C., cited two experiments to show that air is material: (i) A blown bladder resists compression, and (ii) the inside of an inverted drinking-glass when plunged beneath the surface of water remained dry, showing that the presence of air prevented the ingress of the water. These are among the earliest experiments on record.

The terms "atmosphere" and "air" are synonymous and interchangeable, but the word "air" is often used when reference is made to a limited portion of the atmosphere. The word "air" was formerly used in the same general sense that the word "gas" is to-day. Later, the meaning of the word "air" was narrowed to "the atmosphere." The word "atmosphere" is derived from the Greek *ἀτμός* (atmos), vapour; *σφαῖρα* (sphaîra), the sphere. The term "atmosphere" is also applied to the gaseous envelope or medium surrounding any body, whatever be the nature of the gas—air, oxygen, carbon, dioxide, etc. Hence the term "atmospheric air" is often used to emphasize the fact that "air" is the enveloping medium.

The weight of air.—Aristotle (384 B.C.), in spite of some confused ideas on the nature of gases, considered air to be a material substance which possessed weight, because a blown bladder weighed less when empty than when filled with air; but Simplicius, a writer of the sixth century, commenting on Aristotle, said that Ptolemy showed that air has no weight when weighed in air, and that Aristotle's conclusion was vitiated by the condensation of moisture in the bladder derived from the air blown from

the lungs during the inflation of the bladder. Near the beginning of the second century Hero of Alexandria described some experiments to prove that air is a material substance. For instance he said: "If we invert the open end of a vessel, having but one opening, in water, the water does not enter; if a hole be now bored in the upper part of the vessel, water rushes in, and air escapes, as may be felt, for if we place our hand over the orifice, we perceive a rush of wind which is moving air." Galileo Galilei, in 1632, first demonstrated satisfactorily that air possesses weight, and he also made a rough determination of the specific gravity of air by comparing the relative weights of equal volumes of air and water. Refined experiments show that 1,000 c.c. of dry air weigh 1.293 grms. under standard conditions—760 mm. pressure 0° , and at sea-level in latitude 45° . Hence, the specific gravity of air is 0.001293 if water be unity. The specific gravity of air, referred to the standard hydrogen = 2, is 28.75; or with oxygen = 32 as the standard, 28.95.

The death of the hypothesis: "Nature abhors a vacuum."—When a glass cylinder, closed at one end, is filled with water; then closed

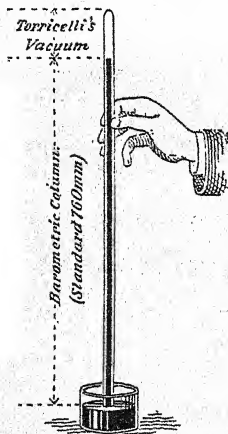


FIG. 22.—Torricelli's Experiment.

at the open end with the hand, turned upside down, and the hand removed while the open end of the cylinder is under water, the water remains in the cylinder. Throughout the Middle Ages, this experiment was explained by the hypothesis: "Nature abhors a vacuum." The rise of water in pump barrels was explained by the same hypothesis. When it was found that water could not be pumped higher than about 34 ft., it followed that the hypothesis required modification, for, as Galilei cynically observed, Nature's horror of a vacuum obviously could only extend to the equivalent of 34 ft. of water.

In 1644, E. Torricelli published an account of an experiment which puzzled the philosophers of the time. A glass tube—about four feet long, and closed at one end—was filled with mercury, the open end was closed with the thumb, and the tube inverted so that when the thumb was removed the open end was immersed in mercury, Fig. 22. No air was allowed to enter the tube during the operation. Instead of the mercury remaining suspended in the tube, as would have been the case with water the column of mercury fell to such an extent that its height above the surface of the mercury in the dish was nearly 30 inches, or 760 mm. Here again, Nature's horror of the vacuum at the top of the tube only extended to the equivalent of 30 inches of mercury. Torricelli abandoned that hypothesis, and concluded, rightly enough, that the column of mercury was maintained by the air pressing on the surface of the mercury in the outer vessel. B. Pascal argued that since mercury is nearly $13\frac{1}{2}$ times as heavy as water, 30 inches of mercury will be equivalent to 34 ft. of water, and he accordingly repeated Torricelli's experiment with a tube 46 ft. long, using red wine instead of mercury. He obtained a column of liquid 34 ft. long. When the experiment was repeated with other liquids, he found, in every

case, that the height of the column was inversely as the density of the liquid. Hence, it was inferred that the height of the column of mercury is a measure of the pressure of the atmosphere, and that fluctuations in the pressure of the air are accompanied by a corresponding fall or rise of the column of mercury. R. Boyle (1665) applied the term *barometer* to Torricelli's instrument—from the Greek *βάρος* (baros), weight; and *μέτρον* (metron), a measure.

In 1647, B. Pascal persuaded M. Périer to repeat Torricelli's experiment at the bottom and at the summit of the mountain Puy-de-Dôme. On September 22, 1648, Perrier wrote: "To our delight and astonishment, the mercury sank lower in the tube the higher up the mountain the tube was carried. This proved that the air presses on the bottom of the mountain more than on the top; and not, as Aristotle and his followers would teach, that Nature has a greater horror of a vacuum at sea-level than at higher altitudes. In a posthumous work, published in 1663, Pascal summarized the arguments which proved conclusively that *all those effects, previously attributed to Nature's horror of a vacuum, are really produced by the pressure, that is, by the weight of the air.* Thus perished the hypothesis: "Nature abhors a vacuum."

The pressure and weight of the air.—The pressure of the air in any given locality varies within comparatively narrow limits. The normal or standard pressure of the atmosphere is equal to the weight of a column of mercury of unit area, and 760 mm. high. This pressure is sometimes called "one atmosphere." It is merely necessary to know the height of the barometric column to know the weight or pressure of the air per unit sectional area. This corresponds with a weight of 1033·3 grms. per square centimetre or 14·7 lbs. per square inch. The word "pressure" is generally used in preference to "weight," because air, like all other fluids, not only presses downwards, but also equally in all other directions.

The density of mercury is 13·596, and in latitude 45° the force of gravity is 980·6 dynes. Hence, a barometric column 76 cm. high will be maintained by a pressure equivalent to $76 \times 13\cdot596 = 1033\cdot3$ grms., or $1033\cdot3 \times 980\cdot6 = 1,013,300$ dynes per sq. cm., or in round numbers 10^6 dynes per sq. cm. This number is called a megabar, and a ten-thousandth part of 10^6 , or 100 dynes per sq. cm. is called a bar. These units are sometimes used for pressure instead of pounds per sq. in., kilograms per sq. mm., etc. A pressure of one megabar is nearly 2 per cent. greater than a kilogram per sq. cm., and 1·3 per cent. less than the atmosphere unit. 1 cm. of mercury = 133·33 bars; and 1 bar = 0·0075 cm. of mercury.

The extent of the atmosphere.—The air gets less and less dense at higher and higher altitudes, and I. Newton (1706) estimated air to be four times rarer at an elevation of about $7\frac{1}{2}$ miles than at sea-level; 1,000,000 times rarer at a height of 76 miles; and 1,000,000,000,000,000 times rarer at an altitude of 228 miles; "and so on." It is therefore impossible to place a limit to the height the atmosphere extends.¹ At a height of 100 to 125 miles, there is sufficient air to offer enough resistance to the passage of meteorites to raise their temperature to incandescence.

¹ G. Johnstone Stoney's memoir: "The Atmosphere of Planets and Satellites" (1897), does not arrive at any definite limit for our atmosphere, nor for the atmosphere of any other planet, but shows that, because the molecules of some gases attain certain high velocities, these gases are able to escape from the atmospheres of the earth and the other planets (see p. 185).

Whatever be the height, the weight of the normal barometric column (per square centimetre of mercury) measures the normal weight of a column of air of the same sectional area and extending from sea-level upwards. B. Pascal (1663) appears to have been the first to calculate the total weight of all the air about the globe. His estimate is approximately equivalent to 4,000,000,000,000 tons.

§ 2. The Influence of Pressure on the Volume of Gases—Boyle's Law.

At the bottom of all cosmic order lies the order of mathematics, the law that twice two is always four.—P. CARUS.

The quantity of matter in a given body is generally determined by weight, but it is often convenient, when the given body is a gas or liquid, to measure the quantity of matter indirectly by volume. Volumetric analysis is based on such measurements; and the analysis of gases is nearly always conducted by volume measurements. The main advantage of measurement by volume is rapid execution; the main advantage of measurement by weight arises from the fact that the result is largely independent of the physical and chemical conditions of the body in question. The weight of a gas is usually so small in comparison with its volume that it is generally possible to determine the quantity of gas more accurately by volume than by weight.

The volume of a gas is very sensitive to changes of pressure. While investigating the relation between the pressure p and the volume v of a gas, Robert Boyle (1661) found "the pressures and expansions," as he expressed it, "to be in reciprocal proportions." In other words, the volume of a gas kept at one uniform temperature varies inversely as the pressure. This is Boyle's Law. E. Mariotte, fourteen years after Boyle's publication, reproduced many of Boyle's results as his own; and, on the Continent, the law is sometimes improperly ascribed to Mariotte.

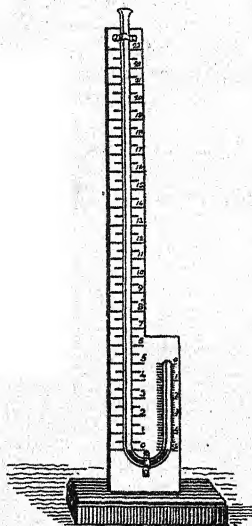


FIG. 23.—Boyle's Law—
High Pressures.

Pressures greater than atmospheric.—The law can be tested in a bent U-tube of uniform bore—Fig. 23—similar to that used by Boyle himself.¹ The shorter leg is hermetically sealed at one end, the end of the longer leg is open. The tube can be graduated by fixing bits of gummed paper at definite distances. Mercury is poured into the longer leg so as to fill the bend and reach to the same height in both legs. It may be necessary to tilt the apparatus a little to expel a few bubbles of air from the shorter leg.* Read the volume of gas confined in the shorter leg. Since the level of the mercury is the same

¹ A great many modified forms of this apparatus have been devised for testing the law. These are intended for quick work in student's "first year" laboratories, or for the lecture table.

in both limbs, it is assumed that the pressure on the surface of the mercury on both sides of the U-tube is the same. There is a pressure of one atmosphere on the mercury in the open leg, hence also there is an equivalent pressure of one atmosphere in the mercury in the closed leg. A pressure of one atmosphere is equivalent to about 30 inches of mercury, or 760 mm. of mercury. In reality, the pressure is equivalent to the height of the mercury barometer at the time of the experiment. Pour a little mercury into the open leg. The gas confined in the shorter leg diminishes in volume. It is easy to prove that no gas has escaped from the shorter leg, and consequently the gas in the shorter leg has been compressed, or is more closely packed than before. In other words, the concentration of the gas per unit volume is increased by the pressure. The difference in the levels of the mercury in the two legs plus the pressure of the atmosphere represents the pressure on the gas in the short leg. More mercury may be poured in the longer leg, and thus a series of numbers are obtained representing the pressure and the volume of the gas in the closed limb. When Boyle had poured sufficient mercury in the longer leg to reduce the volume of the gas in the shorter leg one-half he said, "when we cast our eye on the longer leg, we observed, not without delight and

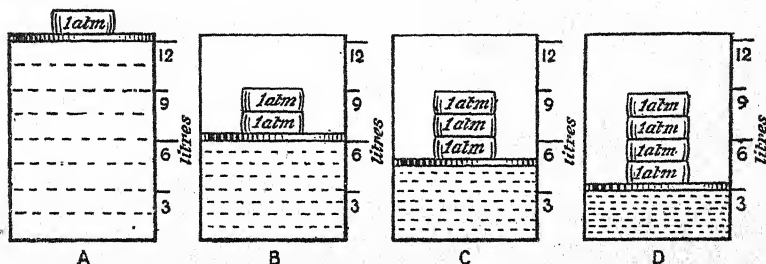


FIG. 24.—Diagrammatic Illustration of Boyle's Law.

satisfaction, that the quicksilver in the longer part was 29 inches higher than in the other." In other words, the volume was diminished one-half when the pressure was doubled by superposing on to the ordinary pressure of the atmosphere the pressure of a column of mercury 29 inches long and equal to the pressure of the atmosphere at the time of the experiment.

To illustrate Boyle's important generalization, imagine 12 litres of a gas confined in a cylinder closed by a gas-tight piston free to slide up and down the cylinder without friction. Suppose further that the gas supports a weight of one atmosphere on the piston, *A*, Fig. 24. If another equal weight be placed upon the piston, *B*, Fig. 24, the gas will be compressed until it occupies a volume of 6 litres; another atmosphere pressure, *C*, Fig. 24, will reduce the volume of the gas to 4 litres; and still another atmosphere pressure, *D*, Fig. 24, will reduce the volume of the gas to 3 litres. Collecting all these results into one table, we see that:

Pressure	1.	2.	3.	4.	6 atmospheres
Volume	12.	6.	4.	3.	2 litres
Product- <i>pv</i>	12.	12.	12.	12.	12

The law of Boyle may therefore be expressed another way: The product

of the pressure and the volume of a gas kept at one uniform temperature is always the same. Or,

$$pv = \text{Constant.}$$

The numerical value of the constant, of course, depends upon what units are selected for representing the pressures and volumes. Pressures may be expressed in atmospheres, millimetres of mercury, pounds per square inch, etc.; and the volumes in litres, cubic centimetres, cubic feet, etc.

Boyle's law assumes yet another guise. If p_1 be the pressure of a gas occupying a volume v_1 , and p the pressure when the volume is v , then, since the products p_1v_1 and $p v$ are equal to the same constant they are equal to one another. Consequently,

$$p v = p_1 v_1$$

If any three of these magnitudes be known, the fourth can be calculated directly.

EXAMPLE.—A eudiometer holds 4.5 litres of gas when the barometer read 755 mm. What will be the volume of the same body of gas when the barometer stands at 760 mm. ? Here, $p_1 = 755$, $v_1 = 4.5$, $p = 760$, hence, $v = 4.47$ litres.

The most common problem is to calculate—"reduce"—the volume of a gas at any observed pressure to the corresponding volume at normal pressure 760 mm. Given 4.5 litres of gas at 755 mm. pressure, there is no need for any formula to calculate the corresponding volume at 760 mm. The pressure, 760 mm., is greater than 755 mm., hence the volume will be less, hence multiply 4.5 by the fraction $\frac{755}{760}$ and the result is 4.47 litres.

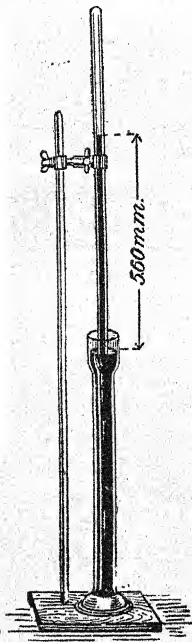


Fig. 25.—Boyle's Law
—Low Pressures.

Pressures less than atmospheric.—Boyle showed that the law holds good at pressures less than atmospheric. Boyle used an arrangement similar in principle to that illustrated in Fig. 25. Some mercury is poured into a narrow tube which is closed at one end and open at the other. The open end is closed by the thumb and inverted in the tall cylinder of mercury. The narrow tube is raised or lowered, and the volume of gas confined in the narrow tube as well as the difference in the levels of the mercury in the narrow and in the wider tube read at the same time. We can recognize the principle of the U-tube, Fig. 23, in this apparatus, Fig. 25. The pressure on the mercury in the wide cylinder is one atmosphere, and the pressure of the gas in the narrow tube is one atmosphere less the pressure of a column of mercury equal to the difference in the level of the mercury in the two tubes.

Measuring the volumes of gases.—This arrangement illustrates a problem which arises very often when the volume of a gas, collected over mercury, is to be measured. If the pressure of the atmosphere is 760 mm., and the difference in the levels of the mercury in the gas jar and in the pneumatic trough is 56 cm., it follows that the

pressure of the gas in the narrow tube is 760 mm. less 560 mm. = 200 mm. Whenever practicable, of course, the mercury inside and outside is brought to the same level before the gas is measured.

Suppose that the confining liquid is water, not mercury. Water is frequently used when the gases are not appreciably soluble in that liquid. Suppose that the external pressure is 760 mm. (barometer), and there is a difference of 10 cm. between the level of the water confining the gas, and the level of the water exposed to the air. The weight of 10 cm. of water is not the same as the weight of 10 cm. of mercury. Mercury is 13.55 times as heavy as water, hence a 10-cm. column of water is equivalent to the weight of a column of mercury $\frac{10}{13.55}$ or 0.74 cm. or 7.4 mm. high. The pressure of the gas is therefore $760 - 7.4 = 752.6$ mm. But water vapour exerts a definite pressure, and a still further reduction must be made if we want the pressure actually due to the gas and not to the mixture of vapour and gas. This will be investigated later.

Test for the equilibrium of gases.—If the gas be confined under such conditions that the product pv at any fixed temperature is not constant, the system will not be in a state of equilibrium. If the piston, referred to in Fig. 24, supports a weight of 6 atm. the gas must occupy a volume of 2 litres; if not, the gas will expand or contract until the product pv satisfies the test. Boyle's law describes the necessary condition for the volume and pressure of a gas to be in a state of equilibrium. In practice there is no such thing as a frictionless piston, and if Boyle's law was to be tested in a real cylinder an allowance would have to be made for the friction of the piston by putting an extra weight on the cylinder. The friction thus corresponds to what J. W. Gibbs (1876) called the *passive resistance* of a system to assume a state of equilibrium. The nature of the passive resistance can here be recognized, but in some cases we feel sure that something analogous retards the movement of a system to the condition called "stable equilibrium," although we know nothing of the character of the passive resistance or the hysteresis—from the Greek *ὑστερέω* (hystereo), I lag behind—which opposes the change.

§ 3. Deviations from Boyle's Law.

Experimentally we do not know of any gas behaving in strict conformity to the law of Boyle; but in the case of many gases, and of nearly all gases at very high temperatures, the deviation from uniformity is very slight.—J. B. STALL.

The pressures used by Boyle extended over a range varying from 3 cm. to 300 cm. of mercury. It is hazardous to infer that because the product pv is constant over a limited range of pressures, it will remain constant for pressures widely different from those actually measured. The method of measurement used by Boyle, excellent for its time, is now considered somewhat crude. Many careful investigations have been made to find if the simple law of Boyle correctly describes the behaviour of gases at pressures far removed from the normal pressure of the atmosphere—76 cm. of mercury. The general results show that no two gases behave precisely in the same way. The deviations for many gases are insignificant. With most gases, the concentration increases more, that is the volume

increases less than Boyle's law describes; and at high pressures, the concentration increases less, that is the volume increases more than Boyle's

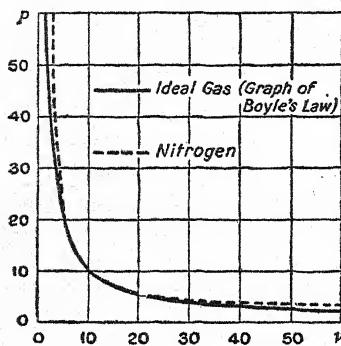


FIG. 26.—Volume : pressure curves.

law indicates. This is illustrated by plotting Boyle's law. Boyle's law, when graphed, furnishes the continuous curve shown in Fig. 26. This curve is a rectangular hyperbola. The deviations with nitrogen from this ideal condition are indicated by the dotted line in the same Fig. 26. According to Boyle's law, the volume of a gas should diminish indefinitely as the pressure is increased, and in time the volume would approach zero, that is absolutely nothing. This is absurd. Pressure can only diminish the space between the molecules and not the actual substance of the molecules. Hence, if b denotes the "volume" occupied by the molecules, the changes in the volume of the gas with variations of pressure will be represented by $p(v - b) = \text{constant}$, not by $p v = \text{constant}$. It does not follow that b represents the actual volume of the space occupied by the matter in the molecules. This subject is taken up later. The effect of the "volume" of the molecules on the compressibility of gas was recognized by D. Bernoulli, 1738; by M. W. Lomonosoff, 1750; by A. Dupré, 1865; and by J. D. van der Waals in 1872.

E. H. Amagat, 1893, showed that while the product $p v$ remains fairly constant at low pressures for many gases, the numerical value of $p v$ changes in a remarkable manner as the pressures increase in magnitude. Amagat's measurements for carbon dioxide show that when

p	1,	50,	100,	125,	150,	200,	500,	1000 atms.
$p v$	1,	0.92,	0.49,	0.31,	0.41,	0.50,	1.02,	1.81

Notice how the product $p v$ at first diminishes in magnitude and then steadily increases. This is brought out very clearly on plotting the numbers. If the products $p v$ were constant for all values of p , we should

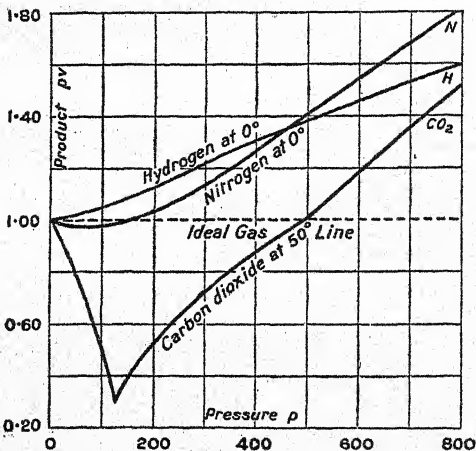


FIG. 27.—Pressure : $p v$ curves (Amagat).

get the straight line, dotted, and marked "ideal gas line" in Fig. 27; with carbon dioxide, however, the curve descends below the line for an ideal gas, and then steadily rises, passing above the ideal gas line when the pressure is nearly 500 atmospheres.

The curves for hydrogen and helium, at ordinary temperatures, do not descend below the ideal gas line, but take a path resembling the hydrogen line shown in Fig. 27. However, even these gases exhibit the same peculiar behaviour at lower temperatures. Thus, with hydrogen at -140° , the product p_v reaches a minimum when the pressure is about 25 atmospheres; at -195° , 45 atmospheres; and at -213° , 51 atmospheres. It was once thought that oxygen behaved in a peculiar abnormal manner at a pressure of about 0.7 mm. of mercury, but some careful measurements by Lord Rayleigh indicate that the statement is probably based upon a mal-observation. To summarize:—

- (1) With small pressures, the product p_v decreases with increasing pressure, showing that the volume with increasing pressure is less than is described by Boyle's law.
- (2) With large pressures, the product p_v increases with increasing pressure, showing that the volume with increasing pressure is greater than is described by Boyle's law.
- (3) All gases, in consequence, show a minimum value for the product p_v . The pressure corresponding with the minimum depends on the nature of the gas and on the temperature. The minimum is less prominent with the more permanent gases¹ than with the more condensable gases.

§ 4. Dalton's Law of Partial Pressures.

Accurate and systematic investigation has brought to light the infinite complexity of nature; the fineness of the dovetailing of every event into many others; the never-ending response of all things to changes in the conditions that encompass them; the universal orderliness of natural occurrences; and the absolute interdependence of cause and effect.—M. M. P. MUIR.

When two gases, which do not act chemically on one another, under the conditions of the experiment, are brought together, the gases mix intimately, by diffusion, so as to form an homogeneous mixture. Furthermore, John Dalton (1802) found that each gas seemed to exert the same pressure as if it occupied the space alone, and the total pressure of the mixture of gases was the sum of the several pressures due to each gaseous component of the mixture. If P be employed to denote the total pressure and p_1 the partial pressure exerted by one of the gases, and p_2 the partial pressure exerted by the other gas, Dalton's discovery means that: $P = p_1 + p_2$. In words, in a mixture of gases which exert no physical or chemical action on one another, each gas exerts the same pressure as if it alone occupied the entire vessel, and the total pressure is the sum of the partial pressures due to each gas. This is Dalton's law of partial pressures. It is independent of Boyle's law, and can be extended to mixtures of any number of gases.

¹ Gases which "obey" Boyle's and Charles' Laws under ordinary atmospheric conditions usually remain gaseous at comparatively low temperatures and are accordingly called *permanent gases*.

EXAMPLE.—Moist hydrogen gas is confined over water under a pressure of 760 mm. of mercury at 10° , the partial pressure of water vapour at that temperature is 9.2 mm. of mercury. Then, from Dalton's law of partial pressures, it follows that the hydrogen gas itself is under a partial pressure equivalent to 760 less 9.2, or 750.8 mm. of mercury.

It is highly probable that the molecules of nearly all gases exert some attractive influence on one another, and the gases will, in consequence of this *physical* action, "deviate" from Dalton's law to an extent dependent upon the magnitude of the intermolecular attraction. Many mixtures of gases show slight, but marked deviations from the law, *e.g.* carbon dioxide and sulphur dioxide; hydrogen with air and with nitrogen, etc. Accordingly, the theoretical results agree more closely with the observed results when an allowance is made for the effect of the attraction of the molecules for one another.

The law has been applied to test if chemical action occurs on mixing certain gases, *e.g.*, to find if any sign of chemical action occurs when nitric oxide (NO) is mixed with nitrogen peroxide (NO_2) resulting in the formation of nitrogen trioxide (N_2O_3). It is assumed that if no chemical combination takes place, the mixture will obey Dalton's law, and conversely. In cases like this, it is assumed that the molecules of the two gases exert neither attractive nor repulsive forces upon one another. If they did, the test might lead to wrong conclusions with respect to chemical action. A slight contraction, for instance, might be evidence of molecular attraction, not of chemical combination.

§ 5. The Laws of Nature.

We must confess that physical laws have greatly fallen off in dignity. No long time ago they were quite commonly described as the Fixed Laws of Nature, and were supposed sufficient in themselves to govern the universe. Now we can only assign to them the humble rank of mere descriptions, often erroneous, of similarities which we believe we have discovered.—J. H. POYNTE.

This is a convenient place to emphasize further the meaning of the term "law" in chemistry. It is of course absurd to say that Dalton's law or Boyle's law *must* be obeyed, implying that these laws are commands imposed upon gases which they are compelled to obey. "Law" is a useful term which the careless sometimes personify. It is employed by scientific men, purely in a metaphorical sense. The term has led to some confusion, and it would be replaced by another word, if we could think of a better. The German equivalent—*Gesetz*, statute—is perhaps worse. As indicated previously, the term "law of nature" is applied to a comprehensive generalization which "methodically and systematically describes certain natural phenomena." The laws of chemical and physical phenomena are collocations of those circumstances which have been found by experiment and observation to accompany all chemical and physical changes included in the statement of the law. The test of the "law" is that the statement holds good without exception. The common meaning attached to the saying, "The exception proves the rule," is wrong, and it is an instance of confusion arising from the double meaning of words. In the old Latin form: *Exceptio probat regulam*, the word "probat" means "tests," just as to-day, "proving wines" means testing them. The

proverb therefore meant that the exception tries, tests, or proves the rule, and if the exception cannot be explained, then the rule breaks down, for the exception disproves the rule. When the exact conditions are set up, the law describes the phenomenon without variableness or shadow of turning. The law is then regarded as an objective power. This power is called a force, and further, the force is said to be the cause of the phenomenon. Thus gravitation is regarded as an attractive force causing one particle to attract every other particle in the universe; chemical affinity is regarded, in this sense, as a selective force which causes certain substances when placed in contact to undergo chemical change. If therefore we find a gas deviating from Boyle's law, or a mixture of gases "disobeying" Dalton's law, the alleged laws may be false, incomplete, or imperfect descriptions, or some perturbing influence is at work which masks the simple phenomena described by these laws. A broken law, said J. H. Poynting, is a false description.

§ 6. The Influence of Temperature on the Volume of Gases—Charles' Law.

According to the schools of philosophy it has been proved that the effect of cold is to make bodies contract, while heat makes them expand.—G. GALILEI (1615).

In 1790, Joseph Priestley concluded "from a very coarse experiment" that "fixed and common air expanded alike with the same degree of heat," and J. L. Gay-Lussac, in 1802,¹ quoted some experiments in support of the broader view: The same rise of temperature produces in all gases the same increase in volume, provided the pressure be kept constant. This law is designated Charles' law, in honour of J. A. C. Charles, who, according to Gay-Lussac, made some crude experiments on the subject fifteen years before Gay-Lussac's publication. Some call this relation "Gay-Lussac's law." It perhaps would be more fittingly called Volta's law, because A. Volta described it in 1793.

The increase in volume which occurs when one litre of nitrogen at 0° is heated in a suitable vessel is shown in the following table (R. Chappius, 1888):—

Temperature 0°.	Volume <i>v</i> litres.	Expansion per litre per degree.
0	1.00000000	
10	1.0367781	0.0036778
20	1.0735396	0.0036776
30	1.1102875	0.0036775
40	1.1470244	0.0036737

The numbers in the last column—called the coefficients of thermal expansion—mean that the volume *v* of a litre of nitrogen, measured at 0°,

¹ John Dalton published an account on some experiments on gases in 1802; but Dalton's statement of the action of heat is not quite right. G. Amontons had an inkling of this law in 1702.

when heated through θ° , can be represented very closely by the expression : $v = (1 + 0.00367\theta)$ litres. In other words, nitrogen increases 0.003676 , or very nearly $\frac{1}{273}$ part of its volume at 0° for every degree rise of temperature. More generally, if v_0 be used to denote the volume of a gas at 0° , we have, instead of the preceding expression,

$$v = v_0 \left(1 + \frac{\theta}{273} \right).$$

This is very nearly true for most of the common gases, and it therefore represents a condition of equilibrium which must be satisfied by the temperature and volume of a gas, under constant pressure, in order that the system may be in stable equilibrium.

While solids and liquids have their own characteristic coefficient of expansion, gases have nearly the same coefficient of thermal expansion. This is the meaning of Charles' law. The coefficients for the gases run something like this :—

Air	0.003665
Hydrogen	0.003667
Carbon dioxide	0.003688

These numbers are close enough to " $\frac{1}{273}$ " for most practical purposes. In general, the more easily a gas is liquefied, the greater the deviation from the constant 0.003665 found for air—witness carbon dioxide.

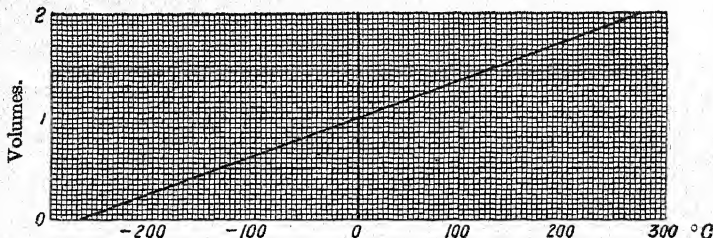


FIG. 28.—Graph of Charles' Law.

By plotting the above equation, we get the curve shown in Fig. 28. If the temperature be less than -273° , the gas would have a negative volume, that is a volume less than nothing! If the temperature be -273° , the gas would occupy no volume! It is impossible to imagine a substance occupying no space, but such is a logical conclusion from Charles' law. Where is the fallacy? Whenever a natural process is represented by mathematical symbols, it is well to remember that the artificial statement often expresses more than actually obtains in nature, because, in the physical world, only changes of a certain kind occur. We must therefore limit the generality of the mathematical expression. Charles' law includes a simplifying assumption. The apparent volume of a gas may be resolved into at least two parts: (1) the "volume" occupied by the molecules of the gas; and (2) the space between the molecules. Although, for the sake of simplicity, we assume v is employed to represent the *total volume* occupied by the gas, in reality v should refer only to the space between the molecules, and in that case, the conclusion that $v = 0$ when

the temperature is -273° involves no absurdity. Moreover the gas would liquefy before the temperature -273° was attained, and the simple gas law of Charles would not then be applicable.

The temperature -273° C. is supposed to be a limiting temperature—the nadir or lowest possible temperature. Hence, it is sometimes called the absolute zero; and temperatures reckoned from this zero are called absolute temperatures. On the absolute scale of temperatures, 0° C. will be 273° abs. If T be employed to denote the temperature on the absolute scale, and θ the temperature on the centigrade scale, we have $T = 273 + \theta$. Hence, we see that if v be the volume of a gas when the absolute temperature is T , and v_1 the volume when the temperature is T_1 , we get, from the preceding equation,

$$\frac{v}{T} = \frac{v_1}{T_1}$$

which is but another way of stating Charles' law.

§ 7. The Combined Influence of Temperature and Pressure on the Volume of a Gas.

According to Boyle's law, the volume varies inversely as the pressure; and according to Charles' law, the volume varies directly as the absolute temperature, then it follows¹ at once that when both temperature and pressure vary, the effect on the volume will be given by the equation

$$pv = RT$$

where R is the constant of proportion—generally called the gas constant. This important relation is sometimes called *Clapeyron's equation*. The same result can be expressed another way: If p , v , and T respectively denote the pressure, volume, and absolute temperature of a gas under one set of conditions, and p_1 , v_1 , and T_1 the pressure, volume, and absolute temperature under another set of conditions, then

$$\frac{pv}{T} = \frac{p_1v_1}{T_1}$$

This formula is used a great deal in calculations involving the variations in the volumes of gases owing to variations in temperature and pressure. For instance in reducing the volume of a gas at any observed temperature and pressure to the corresponding volume at normal temperature and pressure -0° C. and 760 mm. pressure—often represented by “n.p.t.”, or “N.P.T.,” or “S.T.P.” or “S.T., S.P.”

EXAMPLE.—If a gas measures 170 c.c. at a pressure of 735 mm. mercury, and a temperature of 15° , what is the volume of the gas at normal temperature and pressure? Here it is required to find v in the preceding formula, where $p = 760$, $T = 273$, $T_1 = 288$, $v_1 = 170$, and $p_1 = 735$; hence, $v = \frac{760}{735} \times \frac{288}{273} \times 170 = 155.8$ c.c.

§ 8. Deviations from Charles' Law.

Nature abhors the straight line.—R. Ross.

We have already seen that the coefficients of thermal expansion of all gases are only approximately the same. The coefficients for the

¹ It is shown in elementary text-books in algebra that when x varies inversely as y , and x varies directly as z , then $xy = kz$ where k is the constant of proportion.

individual gases differ a little among themselves as indicated above. The variation in the coefficient of thermal expansion at temperatures and pressures, not far removed from normal atmospheric temperatures and pressures, is not very marked, and for regular gas calculations can be ignored. It remains to indicate the variation, if any, in the coefficient of expansion with large variations of temperature and pressure.

1. The influence of pressure.—The coefficient of expansion of most gases is increased by augmenting the pressure of a gas until a maximum value is attained, after that, the coefficient diminishes with increased pressure. For instance, E. H. Amagat (1893) found that the coefficients of expansion of carbon dioxide at temperatures between 50° and 60° assumed the following values:—

Pressure	30,	60,	125,	200,	500,	1000 atm.
Coefficients	0.0069,	0.0085,	0.0410,	0.0085,	0.0033,	0.0018

Carbon dioxide thus shows a marked variation in the coefficient of thermal expansion at high pressures. In agreement with these facts, the coefficient

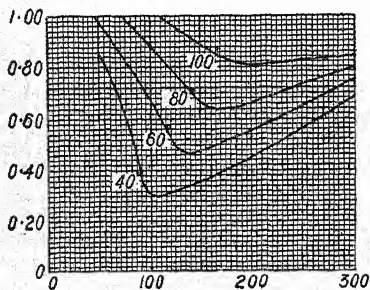


FIG. 29.—Amagat's Curves for Carbon Dioxide.

also diminishes as the pressure is reduced even so low as 0.077 mm. of mercury. The variation is not so marked with gases like nitrogen, oxygen, and hydrogen which are not easily condensed to the liquid condition. The value of p which furnishes the greatest coefficient of thermal expansion is that same value of p which gives the minimum product pv , p. 100. At ordinary temperatures, therefore, hydrogen and helium do not exhibit this variation in the value of

their coefficients of expansion. With these gases, the coefficient of expansion steadily diminishes with increasing pressure.

2. The influence of temperature.—The maximum value for the coefficient of expansion with increasing pressure just indicated becomes less and less as the temperature is raised and finally disappears. So does the minimum value of the product pv become less and less marked as the temperature is raised. The gradual "flattening" of the carbon dioxide curves as the temperature rises from 40° to 100° is brought out very clearly in Fig. 29. We have seen, p. 100, that all gases exhibit the minimum value for pv . The pressure required for the minimum depends on the temperature as well as on the nature of the gas. The minimum is most marked when the gas is near its critical temperature (p. 100). If the temperature is much above the critical temperature, the minimum is very small—with hydrogen the minimum is inappreciable at 0°—Fig. 27. All other gases show a minimum at ordinary temperatures. Hence, Regnault, who discovered this phenomenon, was led to say that hydrogen is a "gaz plus que parfait." But hydrogen also shows the minimum at reduced temperatures as indicated on p. 100.

§ 9. Methods for Measuring the Vapour Densities of Gases and Volatile Liquids and Solids.

Since determinations of molecular weights are usually made to decide between quantities widely different, minor corrections, necessary for exact values, are not required. For instance, if chemical analysis showed that the molecular weight of a compound is some multiple of 20, then a molecular weight of 83, by vapour density methods, indicates that $4 \times 20 = 80$ is the molecular weight of the body. Corrections for the expansion of glass with temperature; the deviation of the weight of a cubic centimetre of water from one gram; the deviation of the vapour from a perfect gas in calculating the volume at normal temperature and pressure, etc., may be neglected.

I. GASES AT ORDINARY TEMPERATURES.

In measuring the relative density of a substance which is gaseous at ordinary temperatures, a large gas balloon of known volume v is counterpoised on the balance by a similar second balloon of approximately the same volume so as to eliminate corrections necessary for the buoyancy of the air. By repeated exhaustions and re-fillings, the balloon is filled with the gas under investigation. The temperature and pressure are respectively t and p . Let w denote the difference between the weights of the full and empty balloon. The volume v_0 of the gas at 0° and 760 mm. pressure is calculated in the ordinary manner, p. 85:

$$v_0 = v \frac{p}{760} \frac{273}{273 + t}; \text{ or } v_0 = 0.3592 \frac{vp}{273 + t} \dots \dots (3)$$

The corrections needed for very exact density determinations need not be here considered, since they do not affect the general principle. From Avogadro's hypothesis the molecular weight of a gas represents the weight of 22.3 litres of the gas if hydrogen = 2 be taken as the standard. Consequently, if w grams of a gas occupy v_0 c.c. at 0° and 760 mm. pressure, 22,300 c.c. will weigh $22,300 w \div v_0$ grms., and this represents the molecular weight, or the relative density of the gas, hydrogen = 2.

EXAMPLE.—585 c.c. of carbon dioxide measured at 18° and 756 mm. pressure, weighed 1.076 gram. What is the molecular weight of the gas? From (3), the 585 c.c. of gas become, at 0° and 760 mm., 546.1 c.c. Hence, the molecular weight is $22,300 \times 1.076 \div 546.1 = 43.9$.

II. SOLIDS AND LIQUIDS WHICH CAN BE VAPORIZED WITHOUT DECOMPOSITION.

A. From the Weight of a Known Volume of the Vapour.

(1) *J. B. Dumas' process* (1826).—A light glass bulb, *A*, Fig. 30, between 100 and 200 c.c. capacity is weighed, and from 6 to 10 grams of the compound under investigation are introduced into the bulb. By means of

a suitable clamp, *D*, the bulb is fixed in a suitable bath, *B*, at a constant temperature 20° to 30° above the boiling-point of the compound under

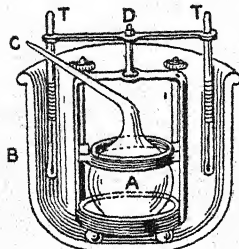


FIG. 30.—Dumas' Vapour Density Apparatus.

investigation. The compound vaporizes, and when its vapour ceases to issue from the neck, *C*, of the bulb, the tube is sealed at *C* by means of a blowpipe with a small flame. The temperature of the bath at the time of sealing is the average between the two thermometers *T*; the barometric pressure is read at the same time. The bulb is then cooled, cleaned, and weighed. The volume of the bulb is now determined by breaking the tip, *C*, of the neck under water or mercury, and weighing the bulb when full of liquid.¹ The difference between the full and empty bulbs gives the amount of liquid in the bulb. The application of the data can be best illustrated by example.

EXAMPLE.—The following data were obtained for vanadium tetrachloride, VCl_4 :

Weight of globe filled with air (9° , 760 mm.)	24.4722 grams
Weight of sealed globe (9° , 760 mm.)	25.0102 grams
Temperature of bath when sealing the globe	215°
Barometer when sealing the globe	762 mm.
Weight of bulb full of water	194 grams

The globe held 194 less 24.4722 = 169.5 grams of water at 9° . This represents very nearly 169.5 c.c. of water, or the capacity of the globe is 169.5 c.c. The

apparent weight of the substance at 9° is 25.0102 — 24.4722 = 0.538 gram. The empty globe was buoyed up, during weighing, by its own bulk of air at 9° and 762 mm., and since 1 c.c. of air at N.P.T. weighs 0.001293 gram, 169.5 c.c. of air at 9° and 762 mm. weigh, at N.P.T. $(0.001293 \times 169.5 \times 273 \times 762) \div (760 \times 282) = 0.213$ gram. This, added to 0.538 gram, gives 0.751 gram, the weight of the vapour in the globe at the time of sealing. The 0.751 gram of vapour occupied 169.5 c.c. at 215° and 762 mm. pressure, or 95.10 c.c. at 0° and 760 mm. pressure. Hence, 22,300 c.c. of vapour at normal temperature and pressure weigh 176.1 grams. This number also represents the molecular weight of vanadium chloride.

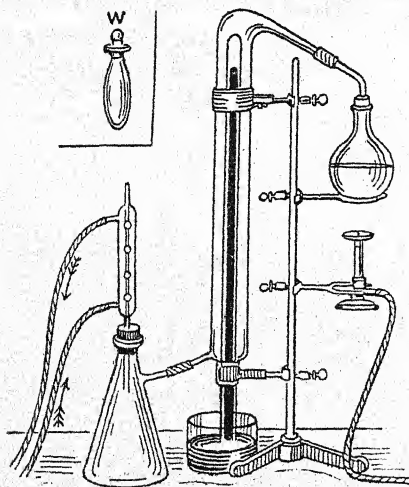


FIG. 31.—Hofmann's Vapour Density Apparatus.

The objection to Dumas' process is the amount of material required to drive out the air from the bulb. This waste is avoided in the two

¹ If the globe contains residual air, a correction must be made. The volume of the vapour will be equal to the volume of the globe, less the volume of the residual air; and the weight of the vapour will be this difference plus the buoyancy of a quantity of air at *t* and *p* of the second weighing, equal to the volume of the vapour.

succeeding methods—Hofmann's and Meyer's processes. By using porcelain or platinum vessels, Dumas' process has been employed for bodies volatilizing at high temperatures.

(2) *A. W. Hofmann's process* (1868).—This is a modification of an earlier method due to J. L. Gay-Lussac (1811). A known weight of the substance in a small stoppered glass bulb, shown on an enlarged scale at *W*, Fig. 31, is taken. The bulb is introduced below a barometer tube filled with mercury, and surrounded with a jacket through which the vapour of a liquid, which boils about 20° above the boiling-point of the compound under investigation, is passing. The bulb ascends to the upper level of the mercury, and the substance is thus vaporized under a reduced pressure. The volume of the vapour is read when everything is in equilibrium; the height of the barometer, and the temperature of the apparatus are also read.

EXAMPLE.—The following data were obtained for carbon tetrachloride, CCl_4

Weight of liquid in bulb	0.3380 grm.
Volume of vapour	109.8 c.c.
Temperature of vapour	99.5°
Barometer	746.9 mm.
Height of mercury in tube	283.4 mm.

The pressure of the vapour, as we shall see very shortly, is the barometric height less the height of the column of mercury in the Hofmann's tube, that is, $746.9 - 283.4 = 463.5$ mm. Hence, 0.3380 gram of vapour at 99.5° and 463.5 mm. pressure occupy 109.8 c.c., and 49.09 c.c. at 0° and 760 mm. Hence, 22,300 c.c. of the vapour at normal temperature and pressure weigh 153.6 grams, and this number represents the molecular weight of carbon tetrachloride.

Hofmann's process is useful when only a small amount of the substance is available for a determination; and for substances which decompose when heated at a temperature in the vicinity of their boiling-point at ordinary atmospheric pressures.

B. From the Volume of a Known Weight of the Vapour.

V. Meyer's process.—V. and C. Meyer (1877) described an elegant and simple method of finding the volume of a vapour by measuring the volume of air displaced by a given weight of the substance vaporized in a suitable vessel. The following is a modification of Meyer's process: A bulb—about 200 c.c. capacity—has a long neck fitted with a side tube leading to a gas-measuring burette, *B*, Fig. 32. An arrangement,

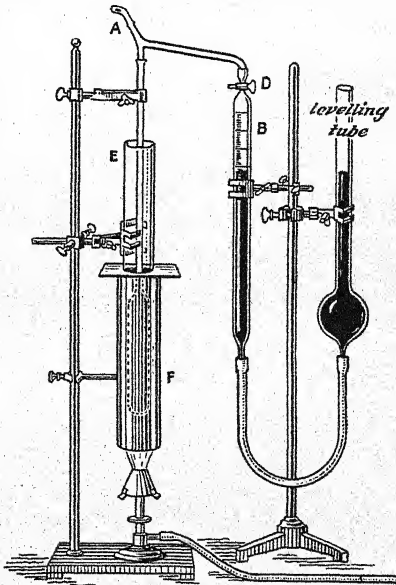


FIG. 32.—Meyer's Vapour Density Apparatus.

A, is fitted to the long neck so that when *A* is turned half a revolution a small stoppered bottle—shown on an enlarged scale at *W*, Fig. 31—can be dropped down the long neck into the heated bulb, which has a little glass wool or asbestos on the bottom, so that the falling bulb will not break the apparatus. A three-way cock, *D*, is turned so as to connect the bulb with the air. The inner chamber is heated by the vapour of a substance placed in the vapour jacket. The substance in the vapour jacket *B* is chosen so that it boils 20° or 30° above the boiling-point of the compound under investigation. The vapour jacket, *E*, is protected in turn by a metal jacket, *F*. When everything is in position: the mercury in the gas burette at zero; the temperature of the vessel constant; and the bulb at *A*, containing a weighed amount of the compound under investigation, is ready to be dropped into the inner chamber, the three-way cock connecting the bulb tube with the atmosphere is turned in order to connect the apparatus with the gas burette. The tube *W* is dropped from *A*. The vapour of the compound displaces its own volume of air, and the displaced air collects in the gas burette. When air has ceased to collect in the gas burette, and the mercury in the levelling and measuring tubes is at the same level, the cock is closed. When the gas burette has had time to attain the temperature of the room, the mercury in the measuring and levelling tubes is again adjusted to the same level, and the volume of air which has collected in the burette is noted. The temperature and the barometer are read at the same time.

EXAMPLE.—The vapour density of water was determined, and the following data were obtained. Xylene, boiling at about 138° , was used in the hot jacket *E*:—

Weight of water in the stoppered tube <i>W</i>	0.0102 gram
Temperature of gas in burette	16.5°
Barometer	703.8 mm.
Volume of gas	16.6 c.c.

The 16.6 c.c. of vapour at 16.5° and 703.8 mm. becomes 14.496 c.c. at 0° and 760 mm. This is the volume of 0.0102 gram of vapour. Hence 22,300 c.c. of the vapour will weigh 15.7 grams. This number represents the molecular weight of water vapour.

If the gas be collected over water instead of over mercury, the correction for the gaseous pressure must be modified. If air contains a per cent. of moisture, and f be the pressure of aqueous vapour at the room temperature, the actual pressure of the confined gas will be $p - (1 - \frac{1}{100}a)f$, where p is the barometric pressure.

Glass vessels are suited for working only at comparatively low temperatures; but vessels of hard porcelain and vitreous silica have been used for temperatures up to about 1500° . Vessels made of iridium lined inside and outside with magnesian cement, and heated in an electric furnace enabled W. Nernst to measure vapour densities as high as 1800° to 2000° .

Questions.

1. State how the volume of a gas is related to temperature and pressure. Describe experiments in illustration of your answer. A gas measures five litres at 20° C., and 770 mm. mercury pressure. What will be its volume at 5° C., and under a pressure of 760 mm. of mercury?—*Aberdeen Univ.*
2. A barometer tube has some air in the space above the mercury. The height of the column is found to be $29\frac{1}{2}$ inches when it ought to be 30 inches and

30 when it ought to be 31. What ought it to be when it reads 29?—*Aberystwyth Univ.*

3. Describe very briefly the principles on which the chief methods for measuring vapour densities are based. What factors determine the choice of one method over others in practice? The vapour density of a substance was determined in two ways. A bulb holding 124 c.c. was sealed up when filled with the vapour at 90°C . and 760 mm. The weight of the vapour was found to be 0.3276 gram. In the second experiment, 0.1 gram of the substance displaced 28 c.c. of air (corrected) in V. Meyer's apparatus. Calculate the vapour density given by the two experiments if 1 litre of hydrogen N.T.P. weighs 0.09 gram.—*Owens Coll.*

4. 0.25 gram of a liquid gave 62 c.c. of vapour measured at 98°C ., bar. 740 mm., the mercury in the measuring tube standing 140 mm. above the mercury in the trough. What was the density of the vapour compared with that of hydrogen?—*New Zealand Univ.*

5. Explain clearly why temperature and pressure observations are necessary when measuring the volume of gases.

6. If a quantity of nitrogen under 900 mm. pressure at 20° occupies a volume of 800 c.c., what volume will it occupy at 100° , under a pressure of 600 mm. pressure?—*Cornell Univ.*

7. 250 c.c. of oxygen at 10° and 756 mm. Find the volume at 0° and 760 mm.—*Mason Science Coll.*

8. If a quantity of hydrogen occupies 500 c.c. in a tube over mercury, the level within the tube being 70 mm. above that, the temperature being 40° , and the barometric pressure 740 mm., what volume will it occupy at S.T.P. (standard temperature and pressure)?

9. Suppose a gas occupies 50 c.c. when confined under normal atmospheric pressure in the measuring vessel A, Fig. 33, with the mercury at the same level inside and outside the measuring vessel, what pressure will be exerted upon the confined gas when the measuring vessel is raised so that the mercury inside is (a) 20 cm. above the level outside the cylinder, B, Fig. 33; and (b) 20 cm. below the outside level of the mercury, C, Fig. 33? What will be the volume of the confined gas in each case? Ansr.: In the tube B, the gas will have a volume of 67.9 c.c.; and in the tube C, 39.6 c.c.

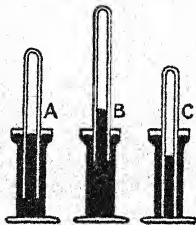


FIG. 33.

10. Oxygen with an atomic weight of 16 has been selected by international agreement as the standard for the comparison of atomic weights. What advantage has the choice of oxygen rather than hydrogen as the standard, and what is the atomic weight of hydrogen in this system? Briefly describe the methods by which the atomic weights are determined, and point out which method is most likely to lead to correct conclusions.—*Cape Univ.*

11. State Boyle's law. A graduated uniform glass tube closed at one end is filled partly with mercury and partly with air; the open end is placed under mercury, and the tube adjusted until the level of the mercury inside is the same as that of the mercury outside. The length of the tube above the mercury is 50 cm. The tube is moved so that the level of the mercury in it is 26 cm. above that of the outside mercury. Find the length of the tube now filled with air when the atmospheric pressure throughout was 76 cm.—*Cape Univ.*

12. In a Boyle's law experiment, the volume of the gas under compression is 100 c.c. when the difference of levels of the mercury in the U-tube is 76 cm. What will be the volume of the gas when the difference of the levels is 152 cm.?—*Sydney Univ.*

13. Explain the difference between atoms and molecules, and give a brief account of the experimental result and of the reasoning by which atomic and molecular weights are determined, choosing an example to illustrate each case.—*New Zealand Univ.*

CHAPTER VI

HYDROGEN

Atomic weight, $H = 1.008$; molecular weight, $H_2 = 2.016$; univalent. Melting-point, -257.14° ; boiling-point, -252.76° ; critical temperature, -241° . Relative vapour density ($H_2 = 2$); (air = 1) 0.0696 ; one litre of hydrogen, at 0° , 760 mm. pressure, and latitude 45° at sea-level, weighs 0.08995 gram (at one time called a *crith*); and one gram occupies 11.117 litres.

§ 1. Hydrogen—Occurrence.

THE element hydrogen occurs free in nature in comparatively small quantities. The atmosphere is said to contain about one volume of hydrogen per 15,000 to 20,000 volumes of air. The proportion of hydrogen in the atmosphere increases when higher altitudes are attained. Hydrogen is also present in volcanic gases; in the gases from the Stassfurt salt beds; and in some meteorites. The presence of hydrogen in natural gas from the oil fields has been denied, although many published analyses of these gases include "hydrogen." The sun's chromosphere shows what appear to be stupendous flames of incandescent hydrogen, in some cases towering over 300,000 miles (M. Fényi, 1892) into space, and 100,000 miles in width (C. A. Young, 1872)—thousands of times larger than the earth on which we live. Spectroscopic observations also show that hydrogen is present in nebulae and certain stars.

Combined hydrogen is common. Water contains one-ninth of its weight of hydrogen. We really know nothing about the hydrogen, as it is combined with oxygen in water. The fact is that when water is decomposed under certain conditions, this proportion of hydrogen is obtained. It is the *façon de parler* to say that the compound "contains" the element, or that the element "occurs" in or is "present" in the compound, when the element can be obtained from the compound by suitable methods of decomposition. Hydrogen, together with oxygen, is one of the chief constituents of animal and vegetable tissue. Hydrogen also is present in nearly all organic compounds, and in many gases—methane, the hydrocarbons, hydrogen sulphide, etc.

Quantitative distribution of the elements.—By comparing a large number of analyses of rocks, etc., F. W. Clarke and H. S. Washington (1922) have tried to estimate the percentage composition—by weight—of

the earth's crust (10 miles deep)—including the ocean and the atmosphere. Their result is:

	Per cent.		Per cent.		Per cent.
Oxygen . . .	49.19	Chlorine . . .	0.228	Nickel . . .	0.030
Silicon . . .	25.71	Phosphorus . . .	0.139	Nitrogen . . .	0.030
Aluminium . . .	7.50	Carbon . . .	0.142	Fluorine . . .	0.030
Iron . . .	4.68	Manganese . . .	0.108	Copper . . .	0.010
Calcium . . .	3.37	Strontium . . .	0.032	Lithium . . .	0.005
Sodium . . .	2.61	Sulphur . . .	0.093	Zinc . . .	0.004
Potassium . . .	2.38	Barium . . .	0.075	Cobalt . . .	0.003
Magnesium . . .	1.94	Chromium . . .	0.062	Lead . . .	0.002
Hydrogen . . .	8.73	Zirconium . . .	0.048	Boron . . .	0.001
Titanium . . .	0.648	Vanadium . . .	0.038	Beryllium . . .	0.001

If we try to get an estimate of the *relative* number of atoms of the different *kinds* of elements distributed in the half-mile crust, the ocean and the atmosphere, Clarke's numbers must be divided by the corresponding atomic weights of the elements. We thus obtain for the *percentage* number of atoms in the half-mile crust:

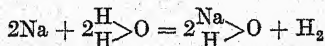
Oxygen . . .	53.81	Sodium . . .	1.72	Potassium . . .	1.02
Hydrogen . . .	16.30	Magnesium . . .	1.61	Carbon . . .	0.27
Silicon . . .	15.87	Calcium . . .	1.40	Titanium . . .	0.16
Aluminium . . .	4.68	Iron . . .	1.29	Chlorine . . .	0.11

This gives a better idea of the relative distribution of the elements from the chemical point of view than the actual weights in the preceding list.

§ 2. The Preparation of Hydrogen.

We have seen in Chapter III. how hydrogen is produced by the action of metals on water or steam. The black oxide of iron produced in the reaction has the empirical formula Fe_3O_4 . The action of steam on metallic iron is represented by the equation: $3\text{Fe} + 4\text{H}_2\text{O} = \text{Fe}_3\text{O}_4 + 4\text{H}_2$. This process is used on a large scale. Porous briquettes of iron are sometimes employed, as in *Lane's process*. These are made from the "spent" pyrites obtained as a by-product in the manufacture of sulphuric acid. In *Bergius' process* the iron is heated with water in a bomb at 300° and 100 atm. pressure. The reaction is represented $\text{Fe} + \text{H}_2\text{O} = \text{FeO} + \text{H}_2$; and if carbon with a little thallium salt as catalytic agent is substituted for the iron: $\text{C} + 2\text{H}_2\text{O} = \text{CO}_2 + 2\text{H}_2$.

The action of sodium on water (p. 66) is represented by the equation:

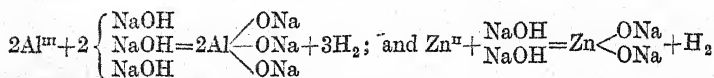


where one atom of hydrogen in the molecule of water is replaced by an atom of sodium, whereby a solution of sodium hydroxide— NaOH —and hydrogen gas are formed. An alloy of sodium with lead—called commercially "hydrone"—generates hydrogen very satisfactorily when in contact with water. The same remark applies to aluminium amalgam—called commercially, "hydrogenite." The same term is applied to a mixture of silicon (25), sodium hydroxide (60), and calcium hydroxide (20), which in contact with water evolves 270–370 litres of hydrogen per kgm. Calcium hydride, CaH_2 , or *hydrolith*, also gives off hydrogen in contact with water:

$\text{CaH}_2 + 2\text{H}_2\text{O} = \text{Ca}(\text{OH})_2 + 2\text{H}_2$. The reactions between iron, zinc, and magnesium with sulphuric or hydrochloric acids (p. 57) are represented: $\text{M}^n + \text{H}_2\text{SO}_4 = \text{MSO}_4 + \text{H}_2$; and $\text{M}^n + 2\text{HCl} = \text{MCl}_2 + \text{H}_2$, where M^n stands in place of the bivalent metal Fe, Zn, or Mg.

In these examples, it will be observed that the process of chemical change results in the substitution of the two atoms of hydrogen in the acid by an equivalent atom of Zn, Fe, or Mg. In the last reaction with zinc and sulphuric acid, for example, zinc sulphate— $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ —can be obtained as indicated previously. Similarly by evaporating the solution remaining after the action of sodium on water, sodium hydroxide— NaOH —is obtained, and with potassium, potassium hydroxide— KOH .

The progress of aeronautics has considerably increased the technical importance of hydrogen. In addition to the methods of preparing hydrogen on a large scale, discussed in Chapter III., a fairly pure gas can be obtained by warming aluminium or zinc with a dilute solution of sodium hydroxide (50 grms. of the hydroxide per 500 c.c. of water). The reaction is represented:



Here it will be observed that the hydrogen atoms in sodium hydroxide are replaced by the respective elements, and sodium aluminate, $\text{Al}(\text{ONa})_3$, is a by-product in the former process; and sodium zincate, $\text{Zn}(\text{ONa})_2$, in the latter case. We see, therefore, that under the stated conditions, an atom of sodium or potassium can displace only one of the two hydrogen atoms in the water molecule— H_2O ; and that zinc or aluminium can displace the other hydrogen atom.

The preparation of hydrogen by the action of sodium hydroxide on the element silicon has been patented as the *silicol process*. The reaction is represented by the equation: $\text{Si} + 2\text{NaOH} + \text{H}_2\text{O} = \text{Na}_2\text{SiO}_3 + 2\text{H}_2$; and also by $\text{Si} + 4\text{NaOH} = \text{Na}_4\text{SiO}_4 + 2\text{H}_2$. But only 80 per cent. of the available hydrogen is obtained before the reaction begins to slow down. If, however, some calcium hydroxide be mixed with the sodium hydroxide, the process appears to be quite satisfactory. 0.8 kilogram of silicon with 1.2 kilogram of alkali will give, in practice, a cubic metre of hydrogen.

Hydrogen is a by-product in the manufacture of chlorine by the electrolysis of aqueous solutions of alkali chlorides: $2\text{NaCl} + 2\text{H}_2\text{O} = 2\text{NaOH} + \text{H}_2 + \text{Cl}_2$. Chlorine is given off at one electrode and hydrogen at the other. The sodium chloride is broken down by the electric current into sodium and chlorine; and the sodium, in contact with water, reacts as indicated above: $2\text{Na} + 2\text{H}_2\text{O} = 2\text{NaOH} + \text{H}_2$.

Hydrogen and oxygen can be manufactured on a commercial scale, where electric power is cheap, by the electrolysis of aqueous solutions—acidulated water, or a dilute solution of sodium hydroxide—under such conditions that the hydrogen and oxygen are kept separate from each other. The gases are separately compressed in steel cylinders—called “bombs”—under a pressure of about 100 or 150 atmospheres.¹

¹ The cylinders of hydrogen are often coloured red to prevent accidentally using a cylinder of hydrogen for one of oxygen.

§ 3. The Hydrogen Equivalent of the Metals.

It is interesting to determine what quantities of the different metals are chemically equivalent to one gram of hydrogen. This can be done by

dissolving the different metals in different acids.¹ A pair of tubes—Hempel's gas burette—*A* and *B*, Fig. 34, are arranged as shown in the diagram. *C* is a flask containing the necessary acid—cold—and a weighed amount of metal. The metal is contained in a test tube in the flask and left there until the necessary adjustments have been made. The levelling tube *B* is raised until the water in *A* is at zero, and on the same level as the water in the levelling tube *B*. The water in the levelling tube should be nearly at the bottom of the tube. The flask is closed with a rubber stopper. Raise or lower *B* until the level of liquid in the tubes is the same. Read the level of the liquid in *A*. The flask is then tilted so that the metal comes in contact

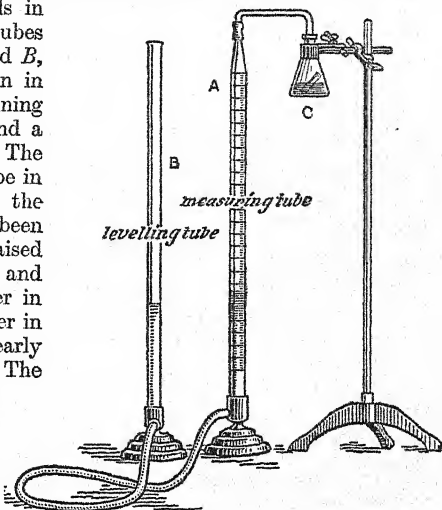


FIG. 34.—The Hydrogen Equivalent of the Metals.

tube is lowered at the same time. When all the metal has dissolved, and the apparatus has had time to cool to the temperature of the room, bring the liquid in the tubes *A* and *B* to the same level by raising or lowering *B*. Read the volume of the gas, the thermometer, and the barometer. Calculate the weight of the hydrogen corresponding with the measured volume of hydrogen, and finally express the result in terms of one gram of hydrogen.

EXAMPLE.—In an experiment by a student, 0.2 gm. of zinc gave 75.5 c.c. of hydrogen, at 15° and 758 mm. pressure. What is the equivalent of zinc? Here 75.5 c.c. at 15° and 758 mm. become 71.3 c.c. at 0° and 760 mm. Since 22.3 litres of hydrogen at 0° and 760 mm. weigh 2.016 grms., 71.3 c.c. of hydrogen will weigh 0.00645 gm. This weight of hydrogen comes from 0.2 gm. of zinc, hence 32.25 grms. of zinc are equivalent to one gm. of hydrogen.

One gram of a given metal will always displace the same amount of hydrogen whatever be the liquid used—*e.g.* aluminium in sodium hydroxide, in sulphuric acid, or in hydrochloric acid; but the amounts

¹ Zinc in dilute sulphuric acid (1 : 7). The strength of dilute acid is often represented by numbers in this way. The numbers mean that one volume of the concentrated acid is mixed with 7 volumes of water. The strength of the commercial acids varies within narrow limits. Magnesium is dissolved in dilute sulphuric acid (1 : 30); in dilute hydrochloric acid (1 : 7); aluminium in warm dilute hydrochloric acid (1 : 3); aluminium in warm dilute sodium hydroxide (50 grms. sodium hydroxide, 500 c.c. of water). There are numerous other ways of doing the experiment besides that described in the text.

furnished by different metals are different. The weight of a metal required to displace one gram of hydrogen is called the hydrogen equivalent, or the chemical equivalent of the metal. The following numbers for the chemical equivalents of three metals are taken from student's note-books:

Zinc, 32.48; magnesium, 12.14; aluminium, 8.96.

These numbers agree closely with those obtained for these metals when referred to the standard oxygen = 8, p. 27.

§ 4. Chemical Affinity.

There are agents in nature able to make the particles of bodies stick together by very strong attractions. And it is the business of experimental philosophy to find them out.—ISAAC NEWTON.

The cause of chemical action has mystified man from the earliest ages, and there is no prospect of an immediate solution. The crucibles, pelicans, and alembics of the working alchemists of the Middle Ages must have demonstrated every day in a thousand different forms that matter seems to be endowed with properties or to possess a kind of occult power in virtue of which two or more dissimilar substances, when brought into contact, give rise to other forms of matter possessing properties quite distinct from the original substances. The process of change is called a chemical reaction. At present, *chemical action can only be referred back to the presence of selective forces, indwelling in the different kinds of matter, which lead certain substances, under certain conditions, to undergo chemical change.* This selective force is called "chemical affinity." Chemists are chary about using the term "chemical affinity" because it suggests that elements, etc., combine or react because of a relationship, kinship, or family tie. The elements fluorine, chlorine, bromine, and iodine, for instance, are said to have a "family relationship" or "natural affinity" for each other. As a matter of fact, elements related in this sense do not usually form stable compounds. On general principles it is well to avoid terms which are not clear and precise in meaning, and which are liable to misunderstanding, because they are not always understood in the same way by everybody.¹ In chemistry, however, the term "affinity" is reserved to connote not a "resemblance" but a tendency of the different kinds of matter to unite with one another. H. Boerhaave used the term in this sense in 1732, and he metaphorically compared the force of affinity with "love, if love be the desire for marriage."² We must allow, said J. B. Dumas (1837), that "there is some truth in this poetic comparison."

It has been said that the term "chemical affinity" is a veil which covers our ignorance in obscure language. It will be observed that to say that "oxygen unites with hydrogen because it has an affinity for it" explains nothing, but restates the fact in different words. So also are statements like "arsenic oxide causes death because it is a poison," "hydrogen burns because it is combustible," "potassium carbonate absorbs moisture from the air because it is deliquescent," and "morphine induces sleep

¹ Some try to evade the difficulty by using other terms: "elective attraction," "chemical activity," "chemical avidity," "chemical energy," etc., but the original term, "chemical affinity," is convenient, provided it is kept in its place.

² . . . *si amor dicendus copulae cupidus.*

because of its soporific qualities." Explanations and definitions of this kind are very rife, and the fault has been given a name—*circular definitions*. We quite recognize with Newton that to tell us that "every species of thing is endowed with an occult specific quality by which it acts and produces manifest effects is to tell us nothing." Hence, the term "affinity" is used by chemists in a metaphorical sense for that peculiar force or form of energy which is the origin of all chemical changes. With this understanding, chemical affinity can be conveniently regarded as "the driving force of a chemical reaction." In 1887, W. E. Ayrton and J. Perry expressed the idea that

$$\text{Reaction velocity} = \frac{\text{Driving force}}{\text{Resistance}}$$

Consequently, if we could measure the chemical resistance offered by substances to undergo chemical change, it would be possible to get a definite and quantitative idea of chemical affinity from measurements on the velocity of a reaction. No real advance can be made in the study of chemical affinity until a method of measurement has been devised.

§ 5. The Measurement of the Affinity between the Acids and the Metals.

I often say that if you can measure that of which you speak, and can express it by a number, you know something of your subject; but if you cannot measure it, your knowledge is meagre and unsatisfactory.—LORD KELVIN.

The relation between the velocity and the driving force of a moving body.—If a ball be sent rolling with a velocity of 20 cm. per second, the force applied to the ball will be twice as great as would be required to make the ball travel with a velocity of 10 cm. per second during the same time. Neglecting friction, the intensities of the two forces are proportional to the velocities which they impart to each mass during the same time.

The relation between the speed of a chemical reaction and affinity.—C. F. Wenzel, 1777, tried to determine the affinities of the metals for different acids by comparing the rates at which the metals liberate gas from acids of different concentration. He found that if an acid of a given concentration dissolves one unit of metal per hour, an acid of half that concentration will take two hours to dissolve the same amount of metal. The velocity of these reactions can be measured by finding the amount of gas liberated per minute, or the amount of acid or of metal consumed, say, every minute. Then, at any given moment:

$$\text{Velocity} = \frac{\text{Amount of gas liberated}}{\text{Time occupied}} = \frac{\text{Acid consumed}}{\text{Time occupied}}$$

The affinity of a metal for an acid depends on the concentration of the acid. In fine, the velocity of the chemical action at any instant is proportional to the concentration of the reacting substances—Wilhelmy's law.

If we start with a solution of such a strength that it contains a gram molecules of acid per unit volume, then at the end of a certain time t , x gram-molecules of the acid per unit volume will have been consumed, and the solution will contain $a - x$ gram-molecules of the acid per unit volume. Hence, the velocity of the reaction will gradually slacken down.

At the beginning of the reaction, the velocity V will be proportional to a ; that is, $V = ka$, where k is a constant;¹ and at the end of the time t , the velocity will be

$$V = k(a - x)$$

Hence, in Wenzel's experiment, when $x = \frac{1}{2}a$, the reaction is only progressing half as fast as at the beginning when $x = 0$. The speed of the reaction with zinc, and with magnesium, at different times, is illustrated in Fig. 35; where the abscissa axis represents time, and the ordinate axis the velocity expressed in any convenient units. The dotted line in the diagram with zinc shows that at the beginning the speed gradually increases, reaches a maximum, and then gradually slackens as described by Wilhelmy's law. In the initial period, called the period of induction, some time is occupied by the reacting substances in inaugurating conditions favourable for chemical action. It is a common feature of chemical

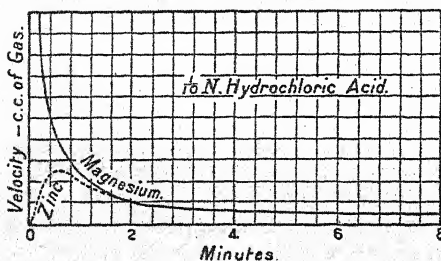


FIG. 35.—The Reduction in the Speed of Chemical Action with Time.

reactions. This period is so short with magnesium in comparison with zinc that the former metal seems to receive a "flying start" as a handicap in the race.

By measuring the rate at which hydrogen is liberated per minute per unit area of the different metals on the same sample of dilute acid, it is possible to get a rough idea of the relative affinities of the

different metals for that particular acid. Experiment shows that with dilute hydrochloric acid, starting with the most vigorous, this order is:

Potassium, sodium, calcium, magnesium, zinc, and iron.

By measuring the rates at which hydrogen is evolved with one metal and different acids of equivalent concentration, we get an idea of the relative affinity of the acids for the given metal.² For instance, acids containing 36.5 grms. of HCl (hydrochloric acid) per litre; 49 grms. of H_2SO_4 (sulphuric acid); and 60 grms. of CH_3COOH (acetic acid) per litre are chemically equivalent to one gram of hydrogen when tested by the methods indicated in Fig. 32. When such acids react with magnesium ribbon (say, 0.05 gm.), the relative affinities appear to be in the order named:

Hydrochloric acid	100
Sulphuric acid	70
Acetic acid	0.5

The first gives off most hydrogen in a given time, the latter least.

¹ Text-books on algebra prove that when x is proportional to y , $x = ky$, where k is a constant.

² The dissolution of metals in acids is not suited for exact measurements because so many disturbing influences are at work—local rise of temperature; bubbles of gas protecting the surface of the metal from attack; variations in the surface of the metal during the action; etc. Still, the conclusion indicated above is in harmony with a great deal of work on a variety of simple reactions.

Returning to the law symbolized in the equation $V = k(a - x)$, with the acids of unit concentration, it follows that the velocity $V = k$. And k has accordingly been called the "affinity constant" of the acid for the metal; k represents the speed of the reaction at the instant when the acid has unit concentration. The result of this discussion shows that the velocity of a chemical reaction is proportional (1) to the "affinity constant" between the reacting substances; and (2) to the concentration of the reacting substances.

§ 6. Opposing Reactions. Guldberg and Waage's Law.

In an aggregate of molecules of any compound, there is an exchange constantly going on between the elements which are contained in it.—A. W. WILLIAMSON (1850).

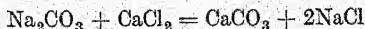
Some of the earlier chemists—*e.g.* Torbern Bergman (1783)—argued that the result of a chemical change must be in favour of that substance with the stronger affinity. Accordingly, "Affinity Tables" were compiled to show the order in which the different substances would displace one another from a given compound: *If A displaces B from one compound, and B displaces C from another compound, the order of the affinity of these three substances is A, B, C.* It was clearly recognized that this method of work does not give a numerical measure of affinity, but it was thought that relative results were obtained. The suggestion is certainly a good trial hypothesis. Let us compare it with the facts.

We have seen that iron can displace hydrogen from its combination with oxygen; hence iron has a stronger affinity than hydrogen for oxygen. Similarly, we have seen that hydrogen can displace iron from its combination with oxygen; consequently, hydrogen has a stronger affinity than iron for oxygen. These two conclusions are contradictory; both cannot be true. Therefore *the affinity hypothesis must be either false, or some powerful perturbing influence must be at work.*

C. L. Berthollet clearly recognized an important disturbing factor, and described it in an heretical but prophetic monograph: *Researches on the laws of affinity*, in 1799. Berthollet noticed large quantities of "trona"—sodium carbonate—on the shores of the natron lakes of Egypt. He suggested that the sodium chloride brought down by the rivers was decomposed by the calcium carbonate present on the banks of these lakes:



Berthollet knew quite well that this reaction is the reverse of that which usually obtains in the laboratory, for sodium carbonate, when added to calcium chloride, precipitates calcium carbonate:



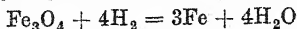
But, added Berthollet, the large masses of calcium carbonate on the banks of these lakes is able to "strengthen" the weak affinity of carbon dioxide for sodium, or of chlorine for calcium. Here Berthollet brings the disturbing factor into bold relief: Chemical action is conditioned not only by affinity but by the relative concentrations of the reacting bodies. Excessive concentration can compensate for a weakness of affinity. A chemical reaction can be reversed by changing the concentrations of the

reacting bodies. We must apply Berthollet's hypothesis to the reaction under consideration—the action of iron on steam.

At the outset, it will be obvious that we have to deal with two opposing reactions: steam reacts with iron to produce iron oxide and hydrogen:



and iron oxide and hydrogen react to produce steam and metallic iron:



Two independent and antagonistic changes take place simultaneously in the system. The result of the change will be determined by the fleetier reaction. When steam is passed over red-hot iron, the hydrogen does not get much chance, it is carried away into the gas jar before it has had time to set up the reverse change. Similarly, when hydrogen is passed over red-hot iron oxide, the steam does not get a chance, for it is carried away from the reduced iron by the stream of hydrogen.

In order to study the affinity relations between these different substances, they should be heated in closed vessels so that the products of the reaction are not whisked away from the seat of the reaction as soon as they are formed. The result is then very curious. It seems as if the reaction stops after a time. At any rate, if the temperature remains constant, no further change can be detected, however long the system be heated. In other words, the system assumes a state of equilibrium. Experiment shows that at 200° the system is in equilibrium when the volume of the steam is to the volume of hydrogen nearly as 1:20. Otherwise expressed, for equilibrium at 200°:

$$\frac{\text{Concentration of steam}}{\text{Concentration of hydrogen}} = \frac{1}{20}$$

If a mixture of one volume of hydrogen and twenty volumes of steam be passed over iron filings or over iron oxide at 200° no apparent change will occur, for the mixture, after passing through the tube at 200°, will have the same composition as when it entered if no secondary actions occur. If more than this amount of hydrogen be present at 200°, some iron oxide will be reduced until the equilibrium ratio 1:0.05 obtains, and conversely, if less than this amount of hydrogen be present, iron oxide will be produced until the equilibrium ratio is obtained. If the temperature be raised, the velocities of the two reactions are altered in such a way that at 440° the concentration of steam will be to that of hydrogen nearly as 1:6, or as 0.17:1; and at 1500°, as 1:1. This means that if equal volumes of steam and hydrogen be passed over iron filings or iron oxide at 1500°, no change in the composition of the gaseous mixture will be perceptible.

Chemical equilibrium is dynamic, not static.—Let us now try to picture what is taking place. Start with metallic iron and steam. At the outset when the reaction is just starting, the velocity of decomposition of the steam will be greatest because the system then contains the greatest amount of reacting substance; and we have seen, p. 117, that "the velocity of chemical action is proportional to the concentration of the substances taking part in the reaction." From this moment, the velocity of the reaction gradually slows down as the concentration of the reacting steam becomes less and less. On the other hand the velocity of the reverse

action will be zero at the commencement, because none of the reacting hydrogen is then present. The speed of the reverse change will become faster and faster as the product of the first-named reaction—hydrogen—accumulates in the system. Ultimately, a point will be reached where the velocities of the two opposing reactions will be equal. The one will be balanced by the other. The reaction will appear to have stopped in spite of the fact that more or less of the original substance still remains untransformed. The system is then in a state of equilibrium. No further change will occur, however long the substances be heated under the same physical conditions of temperature, etc. Chemical changes of this kind are conveniently styled *opposing or balanced reactions*. The idea of a dynamic and not a static equilibrium in such reactions was emphasized by A. W. Williamson about 1850, while studying the action of acids on alcohol. He said: "an exchange is constantly going on between the elements of the molecules of a compound so that each atom of hydrogen in the molecules of HCl present in a drop of hydrochloric acid does not remain quietly in juxtaposition with the atom of chlorine with which it first united, but, on the contrary, is constantly changing places with the other atoms of hydrogen, or, what is the same thing, changing chlorine," and he further adds that when a system appears to be in equilibrium, that condition "is only kept up by the number of exchanges in one direction being *absolutely* the same in each moment of time as those in the opposite direction."

Reversed pointers " \rightleftharpoons " are conventionally used in place of the symbol " $=$ " for opposing reactions, so as to indicate that two reactions are proceeding simultaneously "from right to left" and "from left to right." Accordingly, the reaction under consideration is symbolized:



Opposing reactions are also called *incomplete or reversible reactions* in contradistinction to *irreversible or complete reactions* typified by the action of zinc on sulphuric acid, where the reaction is completed in one direction and is not opposed by a counter reaction.

It is not difficult to see that the *absolute quantities* of steam, hydrogen, iron, and iron oxide, in the reaction under consideration, do not matter. The velocities of the two opposing reactions, and therefore the distribution of the reacting substances, when in equilibrium, is determined by the relative concentrations of the changing substances. This is conveniently expressed by the number of gram-molecules of each present in unit volume. Thus 18 grams of water— H_2O —per litre represents one gram-molecule; 36 grams of water per litre, two gram-molecules; etc. The concentration of a reacting substance is sometimes (inappropriately) called its *active mass*; Berthollet called it *chemical mass*.

If the surface of the iron were doubled, it is true that twice as many molecules of the black oxide, Fe_3O_4 , might be formed in a given time by the decomposition of the steam, but then twice as many molecules of Fe_3O_4 would be decomposed by the hydrogen in the same time. Hence, the amount or the concentration of the solid can have no appreciable influence on the equilibrium; although it may affect the speed at which the state of equilibrium is attained. In studying equilibria in gases and liquids, anything which separates in the solid condition is often

supposed to be thrown out of the reacting system because the state of equilibrium is independent of the concentration of the solid; and a liquid which separates when studying gaseous equilibria, is also supposed to be thrown out of the reacting system. We shall find an analogy in studying the vapour pressure of water in presence of its own liquid. The vapour pressure is independent of the *amount* of liquid water present.

The decomposition and formation of mercuric oxide at different temperatures (p. 22) in a closed vessel is another example of opposing reactions, because the two reactions proceed simultaneously: $2\text{HgO} \rightleftharpoons 2\text{Hg} + \text{O}_2$. The study of this reaction is complicated by the vaporization of the mercury. When the system is in equilibrium, mercury vapour and oxygen are uniting to form mercuric oxide, and mercuric oxide is decomposing into mercury and oxygen. When the speeds of the two reactions are balanced, the gases exert a definite pressure—equilibrium pressure—which measures the concentration of the gases. As the temperature is raised, the speed of decomposition of mercuric oxide is augmented more than the reverse change, and the equilibrium pressure increases as the temperature is raised. One-third of the total pressure of the gases is due to oxygen, and two-thirds to mercury. Hence the course of the reaction can be studied as in the case of iron and steam.

Steam alone is decomposed when heated to a high temperature. The higher the temperature the greater the amount decomposed, or dissociated into its elements: $2\text{H}_2\text{O} \rightleftharpoons 2\text{H}_2 + \text{O}_2$. For instance, W. Nernst and H. von Wartenberg (1906) found:

Temperature	1000°	1500°	2000°	2500°
Amount dissociated	0.00003	0.0221	0.5880	3.98 per cent.

This means that if 100 grms. of steam be heated to 2500°, at atmospheric pressure, the mixture will be in equilibrium when it contains approximately 96 grms. of steam, 3.55 grms. of free oxygen, and 0.45 gm. of free hydrogen. If the temperature be lowered some of the hydrogen and oxygen will recombine; if the temperature be raised more steam will be decomposed. When a substance decomposes with a change in the physical conditions—temperature, pressure, etc.—and the products of decomposition recombine when the original conditions are restored, the process of decomposition is said to be dissociation.

The ideas developed in this section were not so clear to the old workers, not even to Berthollet himself, for Berthollet appears to have confused the incompleteness of certain reactions with the law of multiple proportions. The confusion gave him some strong arguments in the "Berthollet v. Proust Controversy," indicated on p. 28. Proust did not know enough to clarify Berthollet's argument.

Chemical affinity.—To summarize the preceding discussion: chemical affinity is a convenient term for the driving force which causes certain substances to combine together and to remain united with one another.

1. Unlike gravitation, chemical affinity seems to act only when the reacting substances are in contact with one another; or, as it is sometimes expressed: "when the substances are brought within 'insensible' distances of each other."
2. Unlike gravitation, chemical affinity is a selective force, and it seems to act more intensely the more unlike the substances are; or, as it is sometimes expressed: "like reacts with the unlike."

3. The affinity of an element is not only definite as to *kind*, but it is also definite as to the *quantity* of the elements which enter into combination. In this again it differs from gravitational attraction. The quantitative characteristics are described by the "Laws of chemical combination," discussed in Chapter II.
4. The strength of the affinity varies with changes in the conditions of temperature, pressure, light, etc. Gravitational attraction does not appear to be affected by these factors.
5. The velocity of the effect produced by chemical affinity is modified by the relative concentrations—active masses—of the reacting substances.
6. The principles of opposing reactions just outlined are included in Guldberg and Waage's law of mass action—viz. *in balanced or reversible reactions the extent of the chemical change is proportional to the active masses or concentrations of the interacting substances*—so called because the ideas of Berthollet were considerably extended by C. M. Guldberg and P. Waage in an important memoir published in 1864.

§ 7. The Properties of Hydrogen.

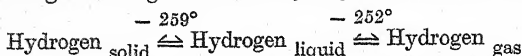
The more salient properties of hydrogen have been discussed on p. 59. Hydrogen is not very soluble in water—100 volumes of water at 0° absorb about 2 volumes of gas, and at 20°, 1·8 volumes of gas. Hydrogen was once used as the standard for the atomic weights because it is the lightest element known. The lightness of hydrogen has been illustrated, Fig. 10, p. 59. A cardboard box or a light glass vessel can be counterpoised, bottom upwards, on a balance. The beam will ascend when hydrogen is poured upwards into the inverted vessel. Soap-bubbles blown with the gas, or collodion balloons filled with the gas, rise to the ceiling very quickly. Hydrogen is accordingly used for filling balloons. Coal gas is used for the same purpose; it is cheaper than hydrogen, though not so buoyant. The lifting power of hydrogen is about 1·2 kgrms. per cubic metre, or about 68½ lb. per 1000 c. ft.

If the molecular weight M (hydrogen 2) of a gas be expressed in ozs., an oz.-molecule of the gas will occupy 22·3 c. ft. at n.p.t.; an oz. of the gas will thus occupy $22·3/M$ c. ft., and a pound of the gas will occupy nearly $357/M$ c. ft., or a c. ft. of the gas will weigh $M/357$ or $0·0028 M$ lb. Again, the lifting power of a volume v of any gas is equal to the difference between the weight of the gas and the weight of an equal volume of air at the same temperature and pressure. The lifting power of a gas of molecular weight M at sea-level is therefore $0·0028v(28 - 98 - M)$ lbs., so that the lifting power of 500,000 c. ft. of hydrogen at n.p.t. is nearly 38,000 lb.

Hydrogen is not poisonous, and animals placed in hydrogen are suffocated for the want of oxygen. When hydrogen is inhaled, the voice becomes shrill—approaching falsetto. The pitch of organ pipes and other wind instruments is raised if a blast of hydrogen be used in place of air. This is in agreement with the fact that the frequency of a note is inversely as the square root of the density of a gas.

Just below the critical temperature, — 241°, a pressure of about 20 atmospheres will liquefy the gas; above the critical temperature, no

pressure, however great, will liquefy the gas. This is one definition of "critical temperature." The liquid is clear and colourless, resembling water, but it has a specific gravity 0.07, and boils at -252° . Hydrogen solidifies when the liquid is evaporated in a partial vacuum. The white solid is crystalline, melts at -259° , and has a specific gravity 0.076. The data concerning the change of state of hydrogen can be symbolized:



The combustible qualities of hydrogen have been indicated on p. 59. The oxyhydrogen flame is one of the hottest gas flames known. When a stick of quicklime is placed at the tip of the flame from a mixture of hydrogen and oxygen burning from a special jet to avoid risk of explosion, the lime does not melt, but it becomes white hot and glows with an intense white light known as "Drummond's light," or the "lime-light" or "calcium light." If zirconia be used in place of lime, the "zircon light" is obtained. The oxyhydrogen flame is also used for the autogenous welding of metals, soldering platinum, and for making vessels of fused quartz.

Catalytic agents.—Although the combustibility of hydrogen is one of its most characteristic properties, perfectly dry hydrogen ignites with difficulty, if at all, when mixed with perfectly dry oxygen. Note, however, that "moisture" is a product of the reaction. Many other combustible substances, if perfectly dried, do not burn when moisture is rigorously excluded. The moisture is here said to act as a catalytic agent—*κατά* (kata), down; *λύω* (lyo), I loosen. Finely divided platinum is also a catalytic agent for a mixture of hydrogen and oxygen, for, in contact with freshly ignited finely divided platinum, or platinized asbestos,¹ these gases can combine at much lower temperatures than they can do in its absence. The platinum suffers no perceptible change during the action.

Hydrogen and oxygen, so far as we can tell, may remain an indefinite time in contact with one another at atmospheric temperatures without showing any sign of chemical action. Some say that the gases do react, but very very slowly. By measuring the diminution in the speed of the reaction from, say, 600° to 500° to 400° to 300° , and assuming that the rate of diminution of the speed of the reaction follows that same law—reduction of the speed one-half per 10° reduction of temperature—it has been stated that no appreciable amount of hydrogen and oxygen will have combined if a mixture of these gases be allowed to stand at ordinary temperatures 1,000,000,000,000 years. The student has the option of accepting or rejecting statements like these. They can neither be proved nor disproved. As P. Duhem has said (1910): "It comes to the same thing experimentally whether we say that the velocity of a reaction is absolutely null, or that it is so small that there is no way of detecting it."

Returning to the catalytic agent—finely divided platinum. If a jet of hydrogen be allowed to impinge on recently ignited but cold platinized asbestos, the mass of platinum becomes hotter and hotter until finally

¹ Asbestos is dipped in a solution of platinum chloride and ignited. A deposit of very finely divided platinum is left on the asbestos, which is then called *platinized asbestos*.

the temperature is sufficient to ignite the hydrogen gas. Similarly, if fresh platinized asbestos be placed in a mixture of hydrogen and oxygen, the gaseous mixture will explode in a short time. The platinum is said to act, not by starting the reaction, but simply by accelerating the immeasurably small speed of combination of the cold gases until they combine at a prodigious rate. The student of chemistry must bear in mind that this is the fashionable view of the function of the catalytic reagent; to-morrow it may be unfashionable. We do not really know how the catalytic agents—moisture and platinized asbestos—act.

"Toy" automatic cigar lighters are made so that by turning the tap of a little hydrogen generator—not unlike the Kipp's apparatus in principle—a jet of hydrogen can be directed on to a piece of spongy platinum when a "light" is desired. The platinum becomes hotter and hotter, and finally ignites the jet of hydrogen. The flame is extinguished by turning the stopcock, and the apparatus is ready for another ignition when the jet of hydrogen is again turned on to the platinum. This is the principle of the self-lighting lamp designed by J. W. Döbereiner in 1822. Impurities in the hydrogen gas, however, appear to "poison" the platinum, for the apparatus soon ceases to be effective.

§ 8. The Action of Hydrogen on the Metals.

Absorption or occlusion.—Some metals—particularly platinum and palladium—absorb relatively large volumes of hydrogen. According to Thomas Graham (1867-9), palladium will absorb 935 times its own volume of hydrogen in cooling from a red heat, and at ordinary temperatures, 276 times its volume. The actual amount absorbed depends upon the physical condition of the metal. According to G. Neumann and F. Streintz (1892), one volume of the following finely divided metals will absorb the following volumes of hydrogen:—

Palladium black	502.3	Nickel	15.6
Platinum sponge	49.3	Copper	4.5
Gold	46.3	Aluminium	2.7
Iron	19.2	Lead	0.1

The hydrogen is given off when the metal is heated, particularly under reduced pressure, and this property of palladium furnishes a useful means of weighing hydrogen gas. It was used by Morley in his work on the combining weights of oxygen and hydrogen (p. 46).

Palladium increases in volume during the absorption, but its general appearance and properties are not much altered, although a considerable amount of heat is evolved during the absorption. Graham called the phenomenon *occlusion* (from *occludo*, I shut up). The gas is said to be occluded by the metal. The phenomenon is now generally called *adsorption*, meaning that the gas *adheres* in some unknown way to the metal. It was once thought that the palladium formed a chemical compound— Pd_2H —with the hydrogen, but this has not been accepted as a full explanation. Graham thought that the gas hydrogen condensed to a solid metal which formed an alloy with the palladium. He gave the name *hydrogenium* to this hypothetical solid to indicate its supposed metallic nature. Solid hydrogen, however, has rather the properties of a non-metal, not a metal—its specific gravity, for instance, it but one-eighth that calculated

for adsorbed hydrogen in palladium. The relation between the adsorbed hydrogen and the metal is not quite clear.

Nascent state.—Hydrogen at the moment of its formation is more chemically active than ordinary hydrogen. For instance, ordinary hydrogen can be passed into an acidified solution of ferric chloride without producing any appreciable change, but if metallic zinc be placed in the solution, the brisk evolution of hydrogen is soon attended with the reduction of the ferric chloride to ferrous chloride. The ferrous chloride gives no reddish coloration with a solution of potassium thiocyanate; the ferric chloride gives a blood-red coloration. Hence, hydrogen in the *nascent condition* at the moment of its birth, can do chemical work which the ordinary gas cannot do. "Hydrogenized palladium" also can do chemical work which ordinary gaseous hydrogen cannot do. It can reduce a solution of ferric chloride to ferrous chloride, for instance:— $\text{FeCl}_3 + \text{H}_{\text{palladium}} = \text{FeCl}_2 + \text{HCl}$. We shall find later on that at least three possible hypotheses are available for explaining this phenomenon! One of these hypotheses regards the hydrogen at the moment of its birth—*nascor*, to be born—as being in the atomic state.

According to I. Langmuir, atomic hydrogen can be obtained by heating filaments of tungsten, palladium, or platinum to a high temperature in a gas of low pressure. The atomic hydrogen so prepared is said to combine with oxygen at room temperature, and with phosphorus to form phosphine, PH_3 . J. J. Thomson, and F. W. Aston showed that triatomic hydrogen, hyzone, H_3 , can exist, and G. L. Wendt and R. S. Landauer (1920) prepared it by the action of α -rays on ordinary hydrogen at atmospheric pressure, and by passing hydrogen through a silent discharge tube (see ozone), or through a vacuum tube at a pressure of 20 to 80 mm. This form of hydrogen is said to react directly with sulphur, arsenic, phosphorus, mercury, and nitrogen, and to reduce potassium permanganate in acid or neutral solutions.

Permeability of the metals to gases.—Hydrogen gas can diffuse through solid platinum, palladium, iron, etc. For instance, if hydrogen be passed through a palladium tube, the gas escapes comparatively quickly through the walls of the metal tube. It has been estimated that about 4 litres of hydrogen can pass through a square metre of palladium per minute; and about half a litre through a square metre of platinum in the same time.¹ The explanation of this action is probably connected with the adsorption of the gas by the metal. This phenomenon has to be taken into consideration in chemical analysis when certain substances are heated over the gas flame in a platinum crucible. Reducing gases, from the coal gas, pass through the walls of the crucible and exert a reducing action on substances being calcined—e.g. manganese oxide.

Hydrides.—Hydrogen forms compounds with most of the non-metals and many of the metals. These compounds are called hydrides. Hydrides of the non-metals are usually stable, while the hydrides of the metals are relatively unstable. Not only is water an oxide of hydrogen, but it can also be regarded as an oxygen hydride. Ammonia— NH_3 —is a nitrogen trihydride; hydrazoic acid— N_3H —is a trinitrogen monohydride; methane— CH_4 —and the hydrocarbons generally are carbon hydrides.

¹ About 130 c.c. of hydrogen will diffuse through a sheet of indiarubber about the same size and thickness in the same time.

Arsenic hydride— AsH_3 , phosphorus hydride— PH_3 , etc., are examples of non-metallic hydrides; while potassium hydride— KH ; sodium hydride— NaH ; ¹ etc., are examples of metallic hydrides. Calcium hydride— CaH_2 —is sold under the name "hydrolith" for making hydrogen. By simply treating hydrolith with water, hydrogen is evolved.

§ 9. The Diffusion of Gases.

Owing to the fact that air is 14 times as heavy as an equal volume of hydrogen under the same physical conditions, if a jar of hydrogen be placed mouth upwards under a jar of air, mouth downwards, most of the hydrogen will flow upwards into the upper cylinder, and air will flow downwards into the lower cylinder—Fig. 36. The action is analogous with what would occur if the lower cylinder contained an oil and the upper cylinder water. The oil and water would change places. The two liquids can be left an indefinite time without mixing, for the two liquids—oil and water—are immiscible. On the contrary, the two gases—hydrogen and air—will spread throughout the two vessels in a short time and in such a way that the two cylinders will enclose a homogeneous mixture of air and hydrogen. The gases are miscible in all proportions. Had the two liquids been soluble in one another—say alcohol and water instead of oil and water—these also would have diffused one into the other so as to form a homogeneous solution of alcohol and water.

The process of diffusion in the case of liquids appears to be very much slower than with gases. The molecules of gases seem to lead a more or less independent existence. This is illustrated by the rapidity with which the molecules of, say, ammonia can travel from one end of a room to the other and affect the sense of smell. In liquids, however, the molecules are much less mobile. This can easily be proved by dropping a small grain of aniline dye into a tumbler of clear still water. The water will be uniformly coloured in a few weeks. The molecules of solid substances have practically lost their mobility. But not all. Carbon laid in contact with pure, hot, solid iron will diffuse into the mass of the metal; gold in contact with lead will, in a few years, diffuse into the lead in appreciable quantities; carbon will diffuse into the body of hot vitrified porcelain; and metallic silver will diffuse into hot glass, staining it yellow.

The transfer of gases in bulk from one vessel to another is an effect of gravitation, whereas diffusion is not an effect of gravitation. Thomas Graham (1832) showed that the speed at which the molecules of a gas can diffuse or travel through thin porous membranes is related to the specific gravity of the gas. For example, hydrogen diffuses nearly four times as fast as oxygen; the relative densities of hydrogen and oxygen are nearly as 1:16; and the relative rates of diffusion of the two gases are nearly as $\sqrt{16} : \sqrt{1}$; i.e. as 4:1. Thus we have Graham's law of diffusion: the relative speeds of diffusion of gases are

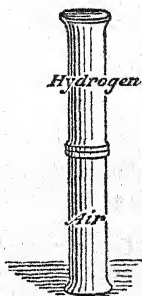


FIG. 36.—
Diffusion Ex-
periment.

¹ The formulæ Na_2H and K_2H , given by L. Troost and P. Hautefeuille (1873), are considered to be erroneous.

inversely proportional to the square roots of their relative densities. Graham measured the speed of diffusion of gases through thin porous plates, and found the numbers indicated in the last column of the subjoined table.¹ The preceding column represents the theoretical numbers calculated on the assumption that the speeds of diffusion are inversely as the relative densities. The observed numbers for the speeds of diffusion agree very closely with those obtained by calculation. There is an interesting application of the law.

TABLE VI.—SPEEDS OF DIFFUSION OF SOME GASES AND GRAHAM'S LAW.

Gas.	Relative density H = 1.	Calculated speed of diffusion from 1	Observed speed of diffusion (Hydrogen = 1)
		$\sqrt{\text{relative density.}}$	
Hydrogen	1	1	1
Methane (CH_4)	8	0.354	0.351
Carbon monoxide (CO)	14	0.267	0.278
Nitrogen	14	0.267	0.265
Oxygen	16	0.250	0.248
Carbon dioxide (CO_2)	22	0.213	0.212

To find the relative density of a gas by comparing its speed of diffusion with that of another gas of known density.—Let D_1 and D_2 represent the relative densities of two gases one of which, D_1 , is known, the other, D_2 , is unknown. Suppose that the relative speeds of diffusion of the two gases V_1 and V_2 are known. Then, it follows from Graham's law: $V_1\sqrt{D_1} = V_2\sqrt{D_2}$. Given any three of these numbers, the fourth can be calculated by arithmetic.

EXAMPLES.—(1) The speeds of diffusion of carbon dioxide and of ozone were found by Soret (1868) to be as 0.29 (V_1) is to 0.271 (V_2). The relative density of carbon dioxide is 22 (D_1) when $H = 1$. What is the relative density of ozone (D_2)? From the preceding relation, it follows that $\sqrt{D_2} = 0.29 \times \sqrt{22} \div 0.271 = 0.29 \times 4.69 \times 3.69$; or $D_2 = (5.02)^2 = 25$ nearly.

(2) A. Ladenberg (1898) found that ozonized air required 367.4 seconds to diffuse under conditions where pure oxygen required 430 seconds: what is the specific gravity of the ozonized air, assuming that the specific gravity of oxygen is unity? Ansr. 1.3689.

Advantage has been taken of the different speeds of diffusion of different gases to devise several pleasing experiments. The simple facts can be illustrated by fixing porous pots² to the ends of bent tubes as shown in the Figs. 37–39. The porous pots contain air. Bring a cylinder of hydrogen over the one porous pot, Fig. 37. Hydrogen diffuses through the walls of the pot faster than the air can diffuse outwards. Consequently, the pressure of the gas inside the porous pot will increase. This is shown by the motion of the coloured liquid in the U-tube away from the porous pot. Before the cylinder of hydrogen was placed over the

¹ I have recalculated Graham's numbers to $H = 1$ instead of air = 1; I have also used whole number approximations for the relative densities.

² Generally used for battery cells. Ordinary clay tobacco pipes with the mouth of the bowl closed with plaster of Paris will serve quite well for the experiment.

porous pot, the air diffused inwards and outwards through the pot at the same rate. Repeating the experiment, Fig. 38, with a cylinder of carbon dioxide instead of hydrogen, the air moves outwards from the porous pot faster than the carbon dioxide can pass inwards. Consequently, there is a reduction in the pressure of the gases in the porous pot. This is shown by the motion of the liquid in the U-tube towards the porous pot.

If the liquid in the leg of the U-tube be connected with a battery and electric bell, and if a wire be fused in the leg of the U-tube so that when the liquid rises electric contact is made, the bell will ring. A device based on this principle has been suggested as an alarm indicator for the escape

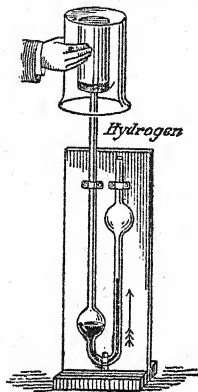


FIG. 37.

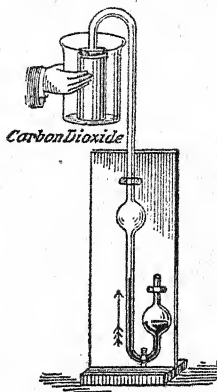


FIG. 38.

Diffusion Experiments.

of coal gas in rooms, or fire-damp in coal mines. These gases, like hydrogen, diffuse through the walls of porous pots faster than the air can escape. The experiments, Fig. 37, can be modified so that the liquid is sprayed from the tube like a miniature fountain. The outward pressure of the gas, Fig. 37, can be made to blow a soap bubble if a film of soap solution be placed across the widened mouth of the tube, Fig. 39.

The ready diffusion of gas through the walls of buildings plays a part in ventilation. Most building materials are porous, and permit the passage of gases through them in both directions. The diffusion does not take place so readily when the walls are saturated with moisture—e.g. new buildings, etc.

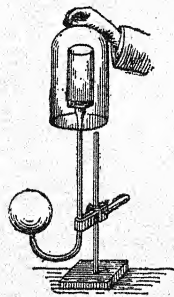


FIG. 39.—Diffusion Experiment.

If a slow current of electrolytic gas, that is, the mixture of hydrogen and oxygen obtained by the electrolysis of water, be allowed to pass through the stem of a "church-warden" clay pipe, and the gas issuing from the pipe be collected in a gas trough, the gas thus collected will no longer explode when brought in contact with a flame. On the contrary, it will rekindle a glowing chip of wood, showing that oxygen is present. In passing through the porous pipe, hydrogen escapes by diffusion through the porous walls of the "clay tube" much more rapidly than the heavier oxygen. This phenomenon—the separation of one gas from another by diffusion—has been called, by Graham, *atmolysis*—*ἀτμός* (atmos), vapour; *λύω* (lyo), I loosen. If a current of steam be passed through a porous tube at a high temperature, and if the porous tube be surrounded by another tube of glazed impervious

porcelain, the water vapour will be dissociated by the heat into hydrogen and oxygen, p. 122; since the hydrogen diffuses much faster than the oxygen, hydrogen will pass from the inner tube into the annular space between the two tubes. The hydrogen may be collected in a suitable gas trough. Similarly, the residual oxygen can be collected from the gaseous steam passing along the inner tube. Of course, the hydrogen is contaminated with more or less oxygen, and the oxygen with more or less hydrogen.

Are the molecules of a gas all alike?—Experiments on atmolysis enable an answer to be returned to this question. If a gas like hydrogen or oxygen be allowed to diffuse through a porous septum, no difference can be detected in the properties of the gas on both sides of the septum. If some of the molecules of, say, hydrogen had a sensibly greater density than the others, it would be possible to produce two kinds of hydrogen by atmolysis. This cannot be done, and hence it is inferred that if there is a difference in the molecules of hydrogen, it is too small to be detected. The alternative is that if there is a definite proportion of two kinds of molecules in equilibrium, equilibrium is restored on both sides of the septum as soon as the two kinds of molecules are separated. W. Ramsay and J. N. Collie (1897) inferred that helium is a simple substance because repeated atmolysis does not alter its density; Lord Rayleigh and W. Ramsay (1895) showed that atmospheric nitrogen is probably a physical mixture of nitrogen with a heavier gas, because repeated atmolysis does alter its density; and F. W. Aston separated neon into two gases with densities 19.9 and 22.1 and hence inferred that neon is a mixture of two gases of different densities.

Questions.

1. Indicate by means of equations four methods of preparing hydrogen. What weight of zinc is required to produce 100 litres of hydrogen gas measured at 27° C. and 680 mm. pressure? ($Zn = 65$, 22.3 litres of oxygen measured at 0° and 760 mm. weigh 32 grams.)—*St. Andrews Univ.*
2. Explain the connection between the terms Equivalent and Atomic weight. 0.100 gram of a metal gave on treatment with a dilute acid 34.2 c.c. of hydrogen measured at N.T.P. Calculate the equivalent of the metal. What further information is required in order to deduce the atomic weight?—*Aberystwyth Univ.*
3. What is meant by "atmolysis"? The specific gravities of air and hydrogen are as 1 : $\frac{1}{14.4}$. How many cubic centimetres of hydrogen will pass through a stucco plug in the same time that 1 c.c. of air will pass through?—*Science and Art Dept.*
4. When 2.331 grams of lead oxide are heated in hydrogen, 2.163 grams of metallic lead are obtained. Suppose the atomic weight of oxygen to be 16, find the quantity of lead combined with 16 parts of oxygen.—*Oxford Senior Locals.*
5. Ten grams of water are (1) decomposed by sodium; (2) passed as steam over red-hot iron; (3) decomposed by an electric current. What volume of gas under standard conditions is produced in each case?—*Queens Coll.*
6. Compare the lifting power of hydrogen of molecular weight 2 with that of helium of molecular weight 4 when used for filling Zeppelin balloons.
7. Calculate the weight of anhydrous zinc sulphate and the volume of hydrogen at 29° and 735 mm. pressure produced by the action of dilute sulphuric acid on 12 grms. of zinc. Hydrated zinc sulphate on heating to 100° loses 37.64 per cent. of water, and on heating to a higher temperature 6.27 per cent. more. Calculate the simplest molecular formula for hydrated zinc sulphate, and find how many molecules of water are lost by the first heating ($Zn = 65$, $S = 32$, $O = 16$, $H = 1$).—*Cape Univ.*
8. State the law of diffusion of gases, and state carefully the ratio of the rate of diffusion of oxygen (molecular weight 32) to that of air (density 14.4).—*Amherst Coll.*

9. A particular gas has a diffusion rate 0.81 referred to air 1.0. Find the density of the gas.—*London Univ.*

10. 120 c.c. of oxygen are mixed with 40 c.c. of hydrogen and the mixture exploded. What volume and weight of gas remain after explosion? Barometer 745 mm., thermometer 12°.—*Owens Coll.*

11. The relation rates of diffusion of oxygen, hydrogen, and nitrogen are respectively 0.96, 3.80, and 1.015. What deduction can be drawn from these figures concerning the density of nitrogen?—*London Univ.*

12. An air-ship with a gas capacity of 100 cubic metres is to be half filled with hydrogen at a temperature of 16° and a pressure of 760 mm. If the hydrogen be obtained by the action of iron on dilute acid, what weight of iron will be required? Why is a balloon never quite filled with gas? $\text{Fe} = 56$. One litre of hydrogen at 0° and 760 mm. weighs 0.9 gram.—*Cambridge Univ.*

CHAPTER VII

THE KINETIC THEORY OF ATOMS AND MOLECULES

§ 1. Matter and Energy.

Substance is like a river in continual flow; the energies undergo constant changes and cause work in infinite variety. There is hardly anything that stands still or remains still.—MARCUS AURELIUS.

MATTER, as perceived by the senses, possesses certain attributes—weight and extension in space—which *appear* to be permanent and essential qualities abiding in *all* known kinds of matter; whereas other properties appear to be secondary and accidental attributes—*e.g.* colour, odour, etc., which are peculiar to *specific* forms of matter. Matter may also be found under different conditions of temperature, electrification, motion, etc., and daily experience teaches us that changes are continually taking place in the conditions of bodies around us. Change of position, change of motion, of temperature, volume, and chemical combination are but a few of the myriad changes associated with bodies in general.

The forms of energy.—It is evident that, in order to keep a grindstone in motion, a certain amount of, say, muscular energy must be expended to overcome the resistance opposed by the air, axle bearings, etc. If a piece of steel be pressed against the stone, the steel soon becomes warm. Exact measurements have shown that the amount of heat produced is proportional to the energy expended in maintaining the motion of the grindstone. Again, in the hot-air motor, heat is employed to set bodies in motion. Heat and mechanical motion are therefore mutually convertible, one into the other. If a vulcanite tire be placed on the grindstone, and the rim be pressed with a piece of flannel, electrification will be induced; but electrical energy can also be readily re-converted back into mechanical motion. Electrical energy, mechanical motion, and heat are thus mutually convertible one into the other. Much of the motive power used in the industrial arts is derived from the chemical action between coal and oxygen in the furnace of a steam-engine. Heat and electricity are also well-known concomitants of chemical action. Hence we infer that heat, electricity, mechanical motion, and chemical action are all different forms of one distinct entity—energy. Observations by Joule and others have shown that any one form of energy can be transformed directly, or by intermediate steps, into any other form. This is the so-called law of transformation of energy. All types of machinery are devices for transforming energy from one form into another.

Energy and work.—As a first approximation, every change in the condition of the bodies around us is supposed to be due to the action of

energy. In other words, energy is that which has the power of changing the conditions of bodies. Whenever a body is changing its condition, there energy is in action. Energy is the cause, change of condition the effect. The action of energy may be resisted. Change can only take place when the restraint is withdrawn or overcome. The action by which energy produces a tendency to change is called a *force*. The word "tendency" here means that the change will take place the moment the restraining influence is withdrawn. Force is thus a manifestation of energy. Whenever resistance is overcome, energy must be expended. Hence energy is sometimes defined as "the power to overcome resistance." Work is said to be performed whenever change takes place in opposition to a force opposing that change. The amount of work done is equal to the quantity of energy transferred. Work is done at the expense of energy.

Work performed = Energy expended.

Consequently, energy is sometimes defined as "the capacity for doing work;" or, as W. Ostwald puts it, energy is work and all else that can be produced from and be converted into work. Two factors are therefore involved in the expenditure of energy: (1) the *magnitude* of the resistance; and (2) the *extent* to which the resistance is overcome. Thus, when a particle moves a certain distance s by the application of a force F , the amount of energy expended, or the work done, is equal to the product Fs .

Energy and matter are inseparable.—We can conveniently describe our knowledge of the material world in terms of two ¹ entities or abstractions: 1. Energy; 2. Matter. It is sometimes convenient to keep these two concepts distinct; although energy and matter are separable only in thought, in reality they are indissolubly joined together. We leave the *metaphysical* ² chemist to deal with matter defined as the unknown cause of known sensations, and answer for himself such questions as: What is matter in and by itself? What is the "*Ding an sich*"? ³ The *working* chemist defines: Matter is that which possesses weight ⁴ and occupies space.⁵ Air, water, glass, copper, etc., are forms of matter; heat, light, electricity, and magnetism are forms of non-matter—energy; colour, odour, etc., are specific properties of particular forms of matter.

Energy, like matter, is indestructible.—So far as accurate measurements are concerned, it is found that when any quantity of one form of energy is made to disappear, an equivalent quantity of another form or

¹ In studying light, etc., it is convenient to add a third—the æther—the hypothetical medium through which the rays of light, etc., are propagated.

² By "metaphysics," I here understand with A. Schopenhauer (*Welt als Wille und Vorstellung*, Berlin, 2, 180, 1854), "every pretended cognition which goes beyond experience, and therefore beyond nature . . . in order to give information about that upon which nature is dependent; popularly expressed, what is behind nature, and makes nature possible."

³ German—"the thing in itself." At the end of our discussion, too, the student might have the uncomfortable feeling that his time had been spent poking into nothing.

⁴ See the footnote, p. 20.

⁵ W. Ostwald (1892): "The more intimately acquainted I became with the properties of energy, the clearer became the proof that matter is nothing but a complex of different factors of energy which possess the property of being reciprocally proportional. The traditional fundamental properties of matter show themselves as modes of expression or factors of energy."

forms of energy appears. No gain or loss of energy has ever been observed in an isolated system. This is the famous law of conservation or persistence of energy. "The transactions of the material universe," said J. C. Maxwell, in that inimitable work *Matter and Motion*, "appear to be conducted, as it were, on a system of credit. Each transaction consists of a transfer of so much credit or energy from one body to another. The act of transfer or payment we call work."

Perpetual motion.—The law of persistence of energy is sometimes called the first law of thermodynamics. It can be expressed another way. No machine can generate energy or do work of itself without consuming at least an equal quantity of "pre-existing" energy. We call this revised statement of the law of persistence of energy, the law of excluded perpetual motion. We can offer no proof of the truth of this law, other than the uncontradicted experience of mankind. We assume that if perpetual motion had been possible it would have been discovered long ago. Of course a similar argument might have been used in 1890 against the existence of a gas like argon in the atmosphere, and the "uncontradicted experience" would have been contradicted four years later. The search for a perpetual motion through centuries of laborious work has been fruitless. It has brought nothing but failure. So great is our faith in the truth of this unproved "law" that a demonstration showing that any supposed process would involve a perpetual motion, *i.e.* the creation or destruction of energy, is considered sufficient proof that the supposed process is impossible. Most scientific societies would refuse to consider seriously papers which violated the assumed law of excluded perpetual motion. It is assumed with M. Faraday that "no hypothesis should be admitted nor any assertion of a fact credited that denies this principle. No view should be incompatible or inconsistent with it."

§ 2. Total, Available, and Potential Energy.

Matter, whatever it is, must be held to be so adorned, furnished, and formed that all virtue, essence, action, and motion may be the natural consequence and emanation thereof.—FRANCIS BACON.

We have no means of measuring the absolute or total amount of energy which a body possesses. Air confined in a closed vessel at atmospheric pressure might appear to possess no energy because it can do no work. But reduce the pressure of the surrounding air, and the air confined in the vessel is capable of performing work. The total energy associated with any body is possibly independent of the external conditions. In the study of natural phenomena, we are only concerned with that portion of the total energy which can be utilized for doing work. This is called the **free or available energy**.

Potential and kinetic energy.—There is an important difference between a stone lying on the ground and a similar stone lying on the table. Both appear alike to be motionless, yet the latter possesses more available energy than the former. For example, the stone, in descending to the ground, could be made to transfer its energy to the mechanism of a clock and do work. The available energy would thus be transformed into mechanical motion. For the same reason, a wound watch-spring possesses more available energy than a similar spring not wound up.

Thus, available energy may be active (*i.e.* kinetic) or passive (*i.e.* latent or potential).

When a marble is rolling along the ground, it has the power, in virtue of that motion, to change the state of another marble with which it might collide. A body, therefore, might possess energy in virtue of its motion. This energy is said to be in a *kinetic* or active condition. It is found that the available kinetic energy K of a body of mass m moving with a velocity V , is $K = \frac{1}{2}mV^2$. This energy may be transformed into heat when the motion of the body is arrested. Potential energy, on the other hand, is said to be "potential to" or "possible to" a body in virtue of its condition with respect to surrounding objects. When a stone is lifted above the ground, the energy expended and the work done depend upon the weight w of the stone and the height h to which the stone is lifted. Consequently, the available potential energy E of the raised stone will be $E = wh$. The meaning is that a measurable quantity of energy is "stored up" or "rendered passive" in some way, and that this same amount of energy can be recovered.¹ For instance, when the stone returns to the ground, it will, in falling, acquire an equivalent amount of kinetic energy. Again, water in an elevated position can do work in virtue of the law that "all liquids will flow to the lowest level that circumstances will permit." Consequently, water at the top of a hill possesses potential energy. A bent spring, a raised hammer, compressed air, and a piece of iron in the vicinity of a magnet, all possess potential energy. Substances which in virtue of their relative condition, or the motions of their molecules, are capable of entering into chemical actions, are also said to possess potential energy. Such is gunpowder, a mixture of metallic zinc and sulphuric acid, etc. The light, heat, sound, and mechanical motion which attend the explosion of gun-cotton are equivalent to the chemical energy stored in the explosive.

§ 3. The Degradation or Dissipation of Energy.

There can be little question that the principle of the dissipation of energy contains implicitly the whole theory of chemical combination.—P. G. TAIT.

The transformations of energy.—Water may be transported from the top of a mountain to the valley beneath in a variety of ways: it may come down in underground channels, rivers, and rain; or in the form of snow, glaciers, or an avalanche. So may energy pass from a state of high to a state of low potential in many and various ways, giving rise to mechanical, thermal, actinic, chemical, electrical, or magnetic phenomena. In reality, the so-called "different forms of energy" correspond with the tendencies which any given system may have to change in particular directions. If there is a tendency for the different parts of a system to come into closer contact, we have gravitation and cohesion; if there is a tendency to an equalization of temperature, thermal energy; and when there is a tendency to undergo transformation into another substance, chemical energy. Hence the definition: a chemical reaction is one mode by which energy can be transferred from one state to another. We have seen many

¹ Note that kinetic and potential energy are here referred to the earth as constant. The suspended stone would have no available energy if it could never fall.

examples of the liberation of energy—heat, light, electricity—during chemical reactions. To avoid the assumption that this energy comes from nothing, it is postulated that the original system contained a definite amount of available energy—chemical energy. As H. Hertz expressed it: In order to explain what is palpably before our eyes, we are compelled to imagine behind the things we see other invisible things, and to search behind the barriers of sense for a secret, hidden accomplice. This hidden factor is conventionally assumed to be potential or latent energy.

If a substance can unite with another, it is said to possess chemical energy, because it can do chemical work; and conversely, substances which cannot combine chemically with other substances have no available chemical energy, for they can do no chemical work. During a chemical reaction, the chemical energy is transformed into an equivalent amount of some other form of energy which is usually, though not always, heat (see p. 33). Hence, the relation between chemical energy and heat (thermal energy) is an important subject, which, for convenience, is called thermo-chemistry. Chemical energy may also be transformed into electrical energy during a chemical reaction. That branch of chemistry which deals with the relation between chemical energy and electricity (electrical energy) is called electro-chemistry. Just as chemical changes which are accompanied by an evolution of heat are called "exo-thermal reactions," so reactions which are accompanied by an evolution of electrical energy have been called "exo-electrical reactions," and conversely for "endo-thermal" and "endo-electrical" reactions.

The degradation of energy.—Just as water will always run down from a high to the lowest level that circumstances will permit, so in all processes with which we are acquainted, every known form of energy at a high potential always runs down to energy at the lowest potential circumstances will permit, and one of the most interesting facts in connection with all natural changes is this constant running down or degradation of energy. Energy becomes less available for doing work. Every change which takes place in nature does so at the cost of a certain amount of available energy. When we inquire whether or not a transformation can take place, the question to be answered is: Will the occurrence involve the degradation of energy? If not, the transformation will not take place under the given conditions. An ancient philosopher has said that "all things are in motion," and we might add that that motion always involves the degradation of energy. The transformation of energy in a given system only ceases when the available energy has run down to the level of its surroundings. The system is then said to be in a state of stable equilibrium.

Metastable equilibrium.—We are, however, very familiar with systems in which the energy has not run down to the level of its surroundings and yet everything appears to be in a state of stable equilibrium. But the stability is only apparent. As a matter of fact, available energy does not always, of itself, run down to the level of its surroundings. For some unknown reason, an influence—conveniently called hysteresis or passive resistance (p. 99)—prevents the initiation of the process of degradation of energy—a preliminary impulse is needed to start the process of degradation of energy. "Passive resistance" is here used as a grouping or classi-

fication term. It explains nothing. Just as the throttle valve of a steam engine must be moved before the engine can start on its journey, so may a preliminary impulse be required to set the process of degradation of energy in motion. The flapping of an eagle's wing may suffice to start an avalanche rolling down the mountain side; with gunpowder the "preliminary impulse" may take the form of heat; with a mixture of hydrogen and oxygen, an electric spark, or the mere presence of spongy platinum; with a mixture of hydrogen and chlorine, a flash of light, or the addition of a piece of charcoal; with fulminate of mercury, a sudden shock; while the addition of a minute crystal will start the process of crystallization of a supercooled solution of sodium thiosulphate. We may thus have a state of metastable, apparent, or false equilibrium, as well as a state of true or stable equilibrium. We naturally inquire: Is there any test to distinguish between states of real and states of apparent equilibrium? We know that if a gas is in equilibrium with regard to volume and pressure it will satisfy the conditions of Boyle's law; volume and temperature, Charles' law, etc. But we have not always such useful tests at our disposal.

§ 4. The Molecular Theory of Matter.

If we would become imbued with the spirit of the new philosophy of chemistry we must begin by believing in molecules.—J. P. COOKE.

For purely chemical reasons, which culminated in Avogadro's hypothesis, we have seen how chemists have been led to invest matter with an imaginary structure which explains, very well, the various transformations which matter undergoes. Matter is supposed to be made up of extremely small discrete particles called molecules. Molecules are the imaginary units which make up matter *en masse*. Molecules are made up of one or more atoms. Atoms are the imaginary units which make up the molecules.

Molecular structure of matter.—Matter must be either a discrete or a continuous medium. Our study of diffusion in solids, liquids, and gases leads us to reject the hypothesis that matter is continuous, for how can two continuous media occupy the same space at the same time? Our study of the compressibility of gases—Boyle's law—leads to the same view. How can a continuous medium on rarefaction (that is, diminution of pressure) expand indefinitely? How can compression diminish the volume of matter itself? If matter be discrete, we can readily answer these queries. Compression involves a closer packing or a crowding together of the molecules by diminishing the space between them. This very explanation was given by Hero of Alexandria 177 B.C. "There are void spaces between the particles of air just as there is air between particles of sand; when the air is compressed, the particles are forced into the vacant spaces, and when the pressure is removed they return to their former position." Conversely, rarefaction involves an increase of the space between the molecules, so that the molecules become less closely packed and less crowded together. If matter be discrete we can also understand how one substance can diffuse into another—hydrogen into air; aniline dye into water, and gold into lead. A mixture of 500 c.c. of water and 500 c.c. of alcohol occupies considerably less than 1000 c.c.; and the difference is greater with a mixture of sulphuric acid and water provided

the temperature of the mixture is the same as that of the separate constituents. We are forced to acknowledge that one substance can be actually penetrated by another, or else assume that the particles of one substance are so disposed that particles of another substance can be accommodated between them much as a scuttle of coal might at the same time accommodate a bucket of sand. As stated, on p. 75, the molecules seem to lead a more or less independent existence, and the space between the molecules furnishes accommodation for the introduction of other particles. A study of the physical and the chemical properties of matter has thus led to one conclusion: Matter is discrete, not continuous; and it is made up of minute particles called molecules. This hypothesis is called the molecular theory of matter.

§ 5. The Kinetic Theory of Molecules.

We must infer that the constituent parts of all bodies are in perpetual motion.
—R. WATSON (1789).

Nature gives no evidence of absolute rest. All matter, so far as we can ascertain, is ever in motion, not merely in masses as with planetary systems, but also molecularly throughout its most intimate structure.—W. R. GROVE.

Are the molecules stationary or in motion? Here again the phenomenon of diffusion has led us to further assume that the molecules are in rapid motion. How could gases diffuse one into the other in such a remarkable way if the molecules were at rest? Diffusion and the fact that a mixture of gases with different specific gravities shows no signs of settling, compel us to assume that the molecules are in a state of incessant motion, and that they are travelling in all directions. In solids, the motion of the particles must be greatly hampered by adjacent molecules. The low compressibility of solids, and the comparatively slow rate at which one solid diffuses into another, shows that the molecules of a solid have a comparatively low mobility. One molecule can only get away from contact with another molecule very very slowly, if at all. The fact that most solids retain their shape for indefinitely long periods, unless prevented by chemical, mechanical, or physical actions, shows that the molecules of solids have a very limited mobility—*e.g.*, some ancient jewellery appears to be the same now as when first engraved. This is sometimes expressed by saying that the motion of molecules is translatory in gases and liquids, and vibratory in the case of solids.

The molecules of a liquid seem to have more freedom than solids. The molecules are sufficiently mobile to allow the liquid to take up quickly the shape of the vessel which contains it. A molecule of a liquid can, in time, travel to any part of the liquid mass. Its course is necessarily slow, because it must be continually abutting against other molecules.

On the other hand, the molecules of a gas seem to lead a more or less independent existence. They appear to be continually moving with a great velocity in sensibly straight lines in all directions. The molecules in their travels must be continually colliding with one another and bombarding the walls of the containing vessel. Thus the molecules continually change their speed and directions.

It is clear that an outward pressure must be exerted on the sides of the vessel every time a molecule strikes the boundary walls. The moving molecules must be perfectly elastic so that after each collision they rebound

with the same velocity as before; otherwise, their momentum would decrease with each collision, and the pressure of a gas would decrease with time, which it does not. Hence, it is inferred that the molecules are in a state of perpetual motion. The preceding assumptions suffice for some important deductions which enable the condition of the molecules of a gas to be inferred with some degree of probability.

§ 6. The Kinetic Theory and Boyle's Law.

What *is* must be studied before what *was* can be inferred. Precedent states remain visionary unless they can be linked to actual and observable conditions.—A. M. CLERKE.

Assume that a closed vessel contains n molecules, and that the ceaseless cannonade of innumerable molecules on the walls of the vessel produces an average pressure, p . Imagine n similar molecules to be squeezed into the same vessel. This will double the number of impacts on the sides of the containing vessel so that the pressure will rise from p to $2p$. The concentration of the gas will also be doubled. This is nothing but another way of stating Boyle's law.

The same result can be obtained another way: Suppose a mass m of gas containing n molecules be confined in a cube with edges each l cm. long, and that the molecules are moving with an average velocity V . Although the molecules travel about in every conceivable direction, it is fair, for purposes of calculation, to consider the molecules are divided into three equal sets with velocities parallel to three adjacent sides of the cube. At any instant, therefore, we assume that $\frac{1}{3}n$ molecules are travelling with a velocity V parallel to any particular edge, and therefore perpendicular to the two corresponding faces of the cube. One molecule moving with a velocity V will take l/V seconds to pass from side to side, and it will therefore strike a side $\frac{1}{2}V/l$ times per second. At each collision with the face of the cube, the velocity of the molecule is reversed in direction so that its momentum changes from mV to $-mV$; that is, its momentum changes $2mV$. The total change of momentum by $\frac{1}{3}n$ molecules striking a side $\frac{1}{2}V/l$ times per second will therefore be the product $\frac{1}{3}V/l \times 2mV \times \frac{1}{3}n$, or $\frac{1}{9}nmV^2/l$. This measures the total force or pressure exerted on one face of the cube. But the total surface of one face of the cube is l^2 . Hence, the total pressure per unit area is $p = \frac{1}{9}nmV^2/l + l^2 = \frac{1}{3}nmV^2/l^3$. But l^3 represents the volume v of the cube. Hence

$$pv = \frac{1}{3}nmV^2.$$

If the number n , the mass m , and the average velocity V do not change, the expression $\frac{1}{3}nmV^2$ is constant, and hence the product pv is constant. This is Boyle's law.

The effect of molecular attraction.—If the molecules have appreciable cohesion, or attraction for one another, they will move in curved, not in straight paths. Doubling the number of particles per unit volume will not then give exactly twice the number of impacts on the boundary walls. When the molecular attraction is marked, the product pv must be *less* than corresponds with Boyle's law. Molecular attraction deflects some of the molecules from the straight path so that they do not strike the walls of the vessel under conditions where they otherwise would. This appears to be the case with carbon dioxide, and most gases which have a smaller apparent volume v , or a smaller value of pv , that is, a greater concentration than corresponds with an increase of pressure as described by Boyle's law. This is illustrated by the downward slope of the pv curves, Fig. 27, for carbon dioxide below 150 atmospheres pressure.

The effect of the size of the molecules.—The small reduction in volume which occurs when a gas is highly concentrated is explained in

the following way: Under great pressures the volume of the molecule becomes comparable in magnitude with the space through which the molecule can move. The volume of the space in which the molecules move is alone reduced by pressure, and therefore only part of the total volume occupied by the gas can be reduced by pressure. Hence, at high pressures the apparent volume, and the product pv appear to be greater than is described by Boyle's law. With hydrogen, for instance, when the pressure is doubled, the volume is not quite halved. The same remark applies to other gases, *e.g.*, carbon dioxide, at great pressures. This is illustrated by the upward course of the curves, Fig. 27.

The kinetic theory and Graham's law.—This relation (p. 127) follows from the above discussion— $pv = \frac{1}{3}nmV^2$. The density of a gas is the mass of unit volume or $mn = D$. Hence considering unit volume, $p = \frac{1}{3}DV^2$. When the pressure is constant, the velocity V will be inversely proportional to the square root of the density D , for $V^2 = \text{constant} \div D$; etc. This is Graham's law.

§ 7. The Kinetic Theory and Charles' Law and Avogadro's Hypothesis.

Nihil enim viveret sine calore.—L. A. SENECA.

Heat is motion which in its strife acts upon the smaller particles of bodies.—

FRANCIS BACON.

That all kinds of fiery bodies have their parts in motion; that heat argues a motion of the internal parts; and that in all extremely hot shining bodies there is a very quick motion that causes light, will be generally granted.—R. HOOKE.

Heat consists in a minute vibratory motion of the particles of bodies.—ISAAC NEWTON.

The kinetic theory and Charles' law.—Heat is generally considered to be a mode of motion of the molecules of matter. The speed of the molecular motion determines the temperature. Each change of temperature, however slight, is supposed to be attended by a corresponding change in the speed of the moving molecules. If the speed increases, the number of molecular impacts on the boundary walls also increases if the volume remains constant; or the volume increases if the pressure remains constant. It is now necessary to introduce an important assumption: *Two gases are in thermal equilibrium when the total kinetic energies of the molecules are the same.* Since real gases—which exert no chemical or physical action on one another and which are under the same conditions of temperature and pressure—can be mixed without change of temperature or pressure, it is assumed that the molecules of equal volumes of two gases at the same temperature and pressure possess the same total kinetic energies. We have seen¹ that the total kinetic energy of a gas is proportional to the product pv . Hence, if the temperature be altered, pressure remaining constant, the kinetic energy (*i.e.* temperature) must alter to the same extent, and hence also the volume. Otherwise expressed, if the pressure remains constant, the same alteration of temperature will alter the volume to the same extent. This is Charles' law.

The kinetic theory and Avogadro's hypothesis.—From what has just been stated, it follows that equal volumes of two gases at the same temperature and pressure have the same value for the product pv . Hence

¹ Magnitudes proportional to the same thing are proportional to one another.

also the total kinetic energy of the one gas will be equal to the total kinetic energy of the other; or the product $n_1 m_1 V_1^2$ for one gas will be equal to the product $n_2 m_2 V_2^2$ for the other. But the average kinetic energy per molecule in the two systems will be equal if the temperature is the same; and hence, $\frac{1}{2} m_1 V_1^2 = \frac{1}{2} m_2 V_2^2$; or, by substitution in the preceding relation, $n_1 = n_2$. This is the symbolic way of saying that equal volumes of two gases under the same physical conditions contain the same number of molecules; that is, Avogadro's hypothesis (see p. 73). It is possible to argue backwards from Avogadro's hypothesis, and deduce the assumption italicized above. The one is dependent on the other. Remember therefore that, contrary to what some enthusiastic writers assert, Avogadro's hypothesis has rendered it necessary to introduce an *unknown and unverifiable assumption*¹ into our reasoning. The kinetic theory should not be quoted as a *proof* that Avogadro's hypothesis is true.

§ 8. Summary of the Kinetic Theory.

The phenomena are our data, and behind them we cannot go except in imagination.—A. SCHOPENHAUER.

We can now summarize the assumptions of the kinetic theory—the term “kinetic,” by the way, is derived from the Greek *κίνηω* (kineo), I move.

(1) *Matter is composed of a finite number of molecules. In gases, the actual volume of the molecules is very small compared with the space not occupied by the molecules. At great pressures, however, the relative size of the molecules must be taken into consideration.*

(2) *The molecules of a gas are in a state of rapid perpetual motion in straight lines. The molecules are continually colliding against the walls of the boundary vessel and against one another.*

(3) *The molecules are perfectly elastic and rebound after a collision without any loss of momentum.*

(4) *The molecules of gases do not always move quite independently of one another since some molecules have a slight attractive force one for the other. This becomes appreciable with increasing concentrations.*

(5) *Two gases are in thermal equilibrium when the average kinetic energies of the molecules of the two gases are the same.*

The kinetic theory and the corresponding molecular theory of liquids and gases have been of great service in helping chemists to form mental pictures of many processes which would be otherwise too difficult to conceive clearly. No one pretends that the picture corresponds with reality, but it has been of great assistance in applying the method of deduction and verification (p. 13). The theory has its faults; at present, it throws no light on many of the properties of gases, while the applications to liquids and solids have scarcely been touched. A great deal of work remains to

¹ According to J. C. Maxwell (1879), “If the system is a gas, or a mixture of gases not acted on by external forces, the theorem that the average kinetic energy for a single molecule is the same for molecules of different gases is not sufficient to establish the condition of equilibrium of temperature between gases of different kinds, such as oxygen or nitrogen, because when the gases are mixed we have no means of ascertaining the temperature of the oxygen and nitrogen separately. We can only ascertain the temperature of the mixture by putting a thermometer in it.”

be done. There is a school of chemists which repudiates the kinetic theory as an exhausted, moribund hypothesis. As a matter of fact, the kinetic molecular theory still promises to live long when we get strong enough to grapple with its many difficulties.

A. D. Risteen (1895) has compared the results of observation with the deductions from the kinetic theory in double columns. The following is modified from his scheme :

Results of Theory.	Results of Observation.
1. The molecules of a gas are all alike.	1. Gases are homogeneous and show no signs of settling, nor can the molecules of any particular gas in general be separated, by diffusion, into gases with different properties. A special case—dissociation—will be discussed later.
2. Molecules are at relatively great distances apart, and in constant motion in straight lines.	2. The compressibility, permeability, and diffusivity of gases are great. The incompressibility of gases at high pressures is supposed to be due to the abnormal crowding of the molecules.
3. In a given mass of molecules, the product pv is proportional to the average kinetic energy per molecule.	3. In a given mass of gas the product pv is proportional to the absolute temperature, etc. This includes the laws of Boyle, Dalton, and Charles.
4. Diffusion.	4. Graham's law.
5. The average <i>kinetic energy</i> is constant for every set of molecules in a mixture of gases.	5. So far as we can tell, the <i>temperature</i> of each constituent of a mixture of gases is the same (see preceding footnote).
6. If two sets of molecules have the same kinetic energy, and the same pressure, they contain the same number of molecules per unit volume.	6. Avogadro's hypothesis, and hence also Gay-Lussac's law. (This is not a result of observation, but it has been inferred independently from purely chemical reasons.)

The first inkling of the idea that the observed properties of matter may be due to motion of its constituent particles has been traced back to Democritus and Lucretius (p. 49). The idea did not develop into a physical hypothesis until R. Hooke (1676), and D. Bernoulli (1738), suggested that gaseous pressure must be due to the impact of the molecules on the sides of the containing vessel. The work of T. Herapath (1821), J. P. Joule (1850), R. Clausius (1857), J. C. Maxwell (1860), and others played important parts in the subsequent development of the theory.

§ 9. Ultramicroscopic Particles.

The sun discovers atoms and makes them dance in his beams.—
D. CULVERWELL.

We are face to face with this extraordinary situation: the molecule has ceased to be a theoretical abstraction—it has become a visible and tangible reality; for we can not only see it, but also “manipulate” it—not, indeed, with our hands, but by means of heat, and electricity, and the air pump.—
E. E. FOURNIER D'ALBE.

In dealing with particles in an extremely fine state of subdivision, the millimetre is an inconveniently large standard of reference. The symbol

μ —pronounced “mu,” and called a micron—is employed for the thousandth of a millimetre, so that $0.001 \text{ mm.} = 10^{-3} \text{ mm.} = 1\mu$; and $m\mu$ —pronounced “double mu” or “millimu,” and called a millimicron—is used for a millionth of a millimetre, such that $0.000001 \text{ mm.} = 10^{-6} \text{ mm.} = 1m\mu$; and $0.001\mu = 1m\mu$. The so-called *Angström's unit* is 10^{-10} metre, and is used in expressing wave-lengths, as well as molecular and atomic dimensions.

Ultramicroscopy.—In practice a good microscope will not clearly resolve particles much smaller than $\frac{1}{2}\mu$ in diameter, and the term *ultramicroscopic particles* is applied to particles smaller than this limit. The ultramicroscopic particles cannot be seen with a powerful microscope illuminated in the ordinary manner, because the light bends round the minute particle and enters the eye just as if the particle did not exist. If the particles be illuminated by a lateral beam of light, their very smallness enables them to scatter the light, so that their presence can be inferred from the fact that each particle is surrounded by visible diffraction rings, just as surely as the presence of smoke indicates fire. The motes dancing in a beam of sunlight would be invisible but for this phenomenon.

Clear solutions, with particles too small to be resolved by the most powerful microscope, appear more or less opalescent when a beam of converging light is focused into the solution. A solution free from these particles would not produce the opalescence, and such a solution is said to be “optically empty.” This is the so-called Tyndall's optical test. Air and gases which are quite free from suspended particles are said to be optically empty because the track of a beam of light therein is invisible. Air can be made optically empty if it is allowed to stand over-night in a glass vessel whose sides are smeared with, say, glycerol. The sensitiveness of J. Tyndall's optical test has been greatly developed by the use of a microscope—called the ultramicroscope—by H. Siedentopf and R. Zsigmondy. In the so-called ultramicroscope an intense beam of light—arc-light, or, better, a beam of bright sunlight—is focused into the liquid under examination, so that the light enters the liquid at right angles to the direction in which it is viewed under the microscope. In one of the earliest experimental methods (1900) of ultramicroscopy, a beam of sunlight was reflected from a mirror, *M*, Fig. 40, through a lens, *L*, and focused in the trough of liquid under examination in the field of a microscope. If transmitted light be used, the eye is dazzled by the profusion of light, and it cannot distinguish the slight differences of brilliancy caused by the diffraction of light by the small particles; just as it is impossible to see the stars by daylight. The later forms of the instrument are more complicated than this, though the principle is the same.

While the opalescence produced by Tyndall's optical test merely shows that a solution contains a number of distinct individual particles in suspension, the ultramicroscope enables the *individual* particles to be detected under conditions where the most powerful microscope would fail to reveal any sign of non-homogeneity. When viewed in the ultramicroscope, the ultramicroscopic particles appear as glittering discs of light with a dim or dark background. A solution may thus appear perfectly homogeneous when viewed under the most powerful microscope, and yet appear distinctly heterogeneous when viewed under the ultramicroscope.

The particles which can be perceived in the ultramicroscope are more

or less approximately the same order of magnitude as the molecules themselves. For instance, ultramicroscopic particles of colloidal gold, $1.7 \mu\mu$, have been measured, and, according to Lobry de Bruyn, the estimated size of a molecule of soluble starch in solution is $5 \mu\mu$; a molecule of chloroform is roughly $0.8 \mu\mu$, according to G. Jäger; and a hydrogen molecule is roughly $0.1 \mu\mu$, according to O. E. Meyer. Hence particles smaller than the complex molecule of soluble starch have been perceived.

The definition of solutions.—Solutions are sometimes defined as “homogeneous mixtures which cannot be separated into their constituent parts by mechanical means (filtration through paper or decantation after settling).” This definition forces us back to the distinction between homogeneous and heterogeneous mixtures; and this, in turn, upon the sensitiveness of the tests for homogeneity. A solution may appear clear and homogeneous; the particles in solution may not be separable by the

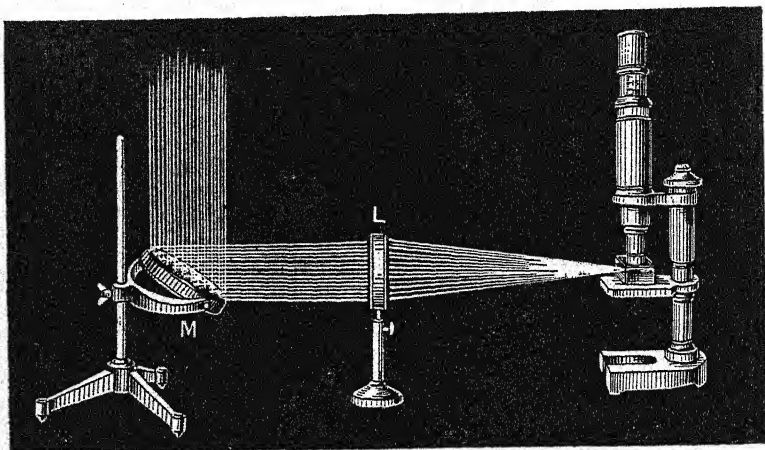


FIG. 40.—Early Form of Apparatus for Ultramicroscopy.

ordinary methods of filtration; and the substance in the solution may remain suspended an indefinite time; and yet when Tyndall's optical test is applied, an opalescence will prove that minute particles are in suspension; and the ultramicroscope will enable the particles to be recognized as distinct individuals. Perfect solutions, said T. Bergman, should be transparent, but there are all possible gradations between liquids carrying rapidly settling particles in suspension; and liquids which carry particles in suspension an indefinite time without settling, and in which the particles are so small that they can only just be perceived by the ultramicroscope. Consequently, if the above definition of a solution be accepted, every time the sensitiveness of the method for detecting non-homogeneity is increased, a certain number of solutions previously classed as homogeneous will probably appear heterogeneous. This difficulty can be partially overcome by arbitrarily restricting the term “solution,” and defining: Solutions are mixtures which appear clear and homogeneous in ordinary day-

light, and which cannot be separated into their constituent parts by ordinary mechanical processes of filtration through paper and decantation after settling. This subject will be taken up again in later chapters.

The Brownian movements.—If water in which a little lycopodium—that is, the spores of the club moss—is suspended be examined under the microscope, the small particles appear to be incessantly vibrating with a slow trembling motion. The phenomenon is named—the *Brownian movement*—after its discovery by R. Brown in 1827. Experiment has shown that the motion cannot be due to convection currents set up by small differences of temperature or pressure, or to any known influence outside the liquid. The cause of the motion must be sought in the liquid itself.

The phenomenon is demonstrated as follows: Rub a fragment of gamboge for a moment on an ordinary 3×1 glass slip, and place a couple of drops of water on the slip where the gamboge has been rubbed. Gently push a cover-glass up to the edge of the gamboge. The brisk motion of the particles can now be readily seen through a $\frac{1}{2}$ " objective and a dark-ground illumination.

The kinetic theory of molecular motion seems to furnish the only admissible explanation of the phenomenon. It is supposed that owing to the perpetual movements of the molecules of the fluid, the moving molecules are continually striking the particles, thus driving them irregularly, to and fro, up and down, in the liquid. As might be expected, the incessant movements become more and more vigorous the smaller the particles. There is a big contrast between the apparently sluggish movements of lycopodium, and the vivacious motions of the ultramicroscopic particles of, say, Faraday's gold. The latter are described by R. Zsigmondy somewhat as follows: "The particles move with astonishing rapidity. A swarm of gnats in a sunbeam will give an idea of the motion. The particles hop, dance, jump, dash together and fly away from one another so that it is difficult to get one's bearings." To this must be added the fact that if the composition of the liquid remains unchanged, the motion in the liquid continues an indefinite time without ceasing.

By studying the movements of the fine particles of gamboge—that is, the dried latex of the *Garcinia morella*—suspended in water; and of extremely fine particles of silver dust obtained by striking an electric arc between silver poles—suspended in air; it has been proved that the distribution of the particles, their velocities, and the frequency of their collision is the same as the kinetic theory assumes to be the case with the particles of a gas. The Brownian movement is thus an expression of the molecular movements usually attributed to the molecules of matter. In fine, the experimental facts go very near towards establishing the validity and essential reality of the molecular kinetic theory as an explanation of the properties of matter.

§ 10. Magnitudes in the Molecular World.

In every explanation of natural phenomena, we are compelled to leave the sphere of sense perceptions, and to pass to things which are not the objects of sense, and are defined only by abstract conceptions.—H. VON HELMHOLTZ.

The above descriptions give no idea of the order of the magnitudes we are dealing with. The kinetic theory allows these magnitudes to be approximately

computed. For air, at atmospheric pressure, and 0° , the calculations furnish roughly:—

1. Diameter of a hydrogen molecule 6×10^{-7} cm.; 24×10^{-9} inch;
0.6 μ .
2. Number of molecules 6×10^{18} per c.c.; 24×10^{19} per
cu. in.
3. Distance traversed between two collisions 0.00001 cm.; 0.000004 in.;
10 μ .
4. Collisions per second 5000,000,000.
5. Velocity per second 45,000 cm.; 1500 feet.
6. Average spacing of molecules in air 10^{-5} cm. or μ .
7. Average spacing of molecules in liquids 10^{-4} cm. or 10 μ .

Most of these numbers convey no meaning to the mind because they are utterly beyond the range of our comprehension. The following considerations will serve to emphasize our inability to form a clear concept of the "scale of magnitudes" in the "world of molecules." *First*, A normal human eye, at a distance of 10 inches, can see objects $\frac{1}{100}$ inch in diameter; with a good microscope objects not much smaller than $\frac{1}{10000}$ inch in diameter can be clearly seen, but this is nearly 5000 times the magnitude of the molecule of an element. It would take about 40,000,000 molecules, touching one another, to make a row an inch long. *Second*, If all the molecules in a cubic inch of a gas were laid in a row, touching one another, they would form a line about 35,000,000 miles long, and this line would extend more than 1000 times round the earth. *Third*, If the gas were magnified on such a scale that a molecule was an inch in diameter, each cubic foot would contain about one molecule, and a molecule would then travel about 100 feet before it collided with another. *Fourth*, It would take about 53 years, counting at the rate of three per second, 24 hours a day, to count the number of collisions—5000,000,000—made by a molecule with its fellows every second. *Fifth*, A molecule travels at the rate of nearly a quarter of a mile per second.

§ 11. Brown's Kinetic Theory of the Atoms.

The curve described by a single atom is as fixed as the path of a planet, and between the two cases no other difference exists, save that resulting from our ignorance.—L. MEYER.

There are countless worlds in countless heavens each revolving about its sun.—G. BRUNO.

Like the planets and satellites of the solar system, the atoms of a molecule are in harmonious stable movement.—D. MENDELÉEFF.

1. Do the atoms of the molecules of a compound retain their individuality? It may be quite true that the properties of a compound are mainly determined by the character of the constituent elements, yet, it is not to be supposed that there is necessarily any resemblance between the properties of the elements and of their compounds. For instance, the properties of a molecule of, say, water are very different from the properties of the constituents hydrogen and oxygen. Although the atoms of a compound molecule do not enjoy a separate external existence, yet, *within* the molecule, the atoms are probably distinct, self-contained, and self-existent; as Lucretius would have expressed it, they are "strong in their solid singleness." The individual properties of the atoms, however, are not always unrecognizable in the properties of the molecules of their compounds. J. Larmor (1908) has well said:

It becomes increasingly difficult to resist the simple view that chemical combination involves an arrangement of the atoms alongside each other under steady cohesive affinity, *the properties of each atom being somewhat modified, though not essentially, by the attachment of the others*; and that the space formulæ of chemistry have more than an analogical significance. The many instances

in which the physical properties of the compound molecule can be calculated additively ¹ with tolerable approximation from those of the constituent atoms, are difficult to explain otherwise.

While iron is intensely magnetic alone and in some of its combinations, in others it appears to be non-magnetic, and in still others, diamagnetic; the atoms of oxygen are magnetic when associated together, but non-magnetic when compounded with many other elements. These facts suggest that the atoms of oxygen lose their individuality in combination. The atom does preserve a kind of individuality throughout the whole of its existence; and when associated with other atoms it may change its habits but not its nature. The change in its habits depends entirely on its associates. This is evidenced by (1) the weight of an atom remains intact throughout the whole of its chemical migrations whatever be its associated partner; (2) each atom emits a peculiar type of X-ray which can neither be changed nor disguised by association with other atoms; and (3) the absorption of X-rays and cathode rays is an atomic property, each kind of atom having its own specific absorptive power which is independent of the nature of the partners with which it may be associated.

2. Are the atoms of a molecule at sensible distances apart? It is sometimes asserted that the atoms are at "insensible" distances apart, and that the atoms of a molecule are accordingly very close together. These statements have given rise to a misconception, for if the size of the atom be taken as a standard of reference it is probable that in the molecule the distances of the atoms from one another are comparatively great.

3. Are the atoms of a molecule at rest or in motion? Some circumstantial evidence bearing on this question will be discussed towards the end of this book. S. Brown (1843) and D. Mendeléeff (1868),² like many previous chemical philosophers, picture a complex molecule as a kind of miniature solar system with the atoms whirling about one another at great speeds. Like the planets and the satellites, the atoms are supposed to be "endowed with an everlasting motion." The atoms are further supposed to be held in position, and to move in definite orbits owing to their attraction for one another, just as the planets and satellites move in definite orbits owing to the action of gravitational forces. Mendeléeff, like Brown, is an enthusiast; he says:

Chemically, the atoms may be likened to the heavenly bodies, the stars, sun, planets, satellites, etc. The building up of the molecules from atoms, and of substances from molecules is then conceived to resemble the building up of systems, such as the solar system, or that of twin stars, or constellations from individual bodies. This is not a simple play of words in modern chemistry, nor a mere analogy, but a reality which directs the course of all chemical research, analysis, and synthesis.

According to this view, a two-atom molecule of hydrogen, H_2 ; iodine, I_2 ; or oxygen, O_2 , might be depicted as a binary star—that is, as a pair

¹ ADDITIVE PROPERTIES.—Additive properties depend on the nature of the atoms in a molecule. Each atom exerts its own specific influence whatever be its state of combination. Numerous examples will occur later—molecular weight, specific heat, crystalline form, etc.; see also pp. 698 and 699.

² T. Bergman and G. L. L. de Buffon both abandoned the theory as impracticable because they failed to realize the importance of the second question indicated above.

of stars. Each atom in the molecule rapidly revolves about the other in a regular orbit—an attempt has been made to illustrate the idea in Fig. 125; a molecule of water, H_2O , would be represented by three atoms revolving in a similar manner; and a molecule of sulphuric acid might be depicted as a complex system with a central revolving sulphur atom around which the other atoms whirl in definite orbits. First would come two oxygen atoms representing the nucleus SO_2 ; outside these would encircle two oxygen atoms each with a revolving hydrogen atom as satellite. The imaginary picture so obtained would be a kinetic model of the molecule $(\text{HO})_2 = \text{S} = \text{O}_2$, or H_2SO_4 . The chemist determines the constitution of these tiny systems by a process which Martin has compared with the plucking of, say, the earth and moon from the solar system, or by replacing one planet by another and observing the disturbing effects of the transposition on the whole system. The case of sulphuric acid will be discussed later. A kind of orrery would therefore give a better idea of the structure of a molecule than the crude plane formulæ usually employed. By this analogy, the planets Mercury and Venus represent single atoms; the Earth, Jupiter, and Saturn with their moons represent radicles—each composed of several distinct atoms so as to form a small sub-system complete in itself. All these individuals and sub-systems are linked to one another so as to form a balanced or stable molecular system, in some respects analogous with the solar system.

4. Valency.—Supposing Brown's and Mendeleëff's speculations were to be established by unassailable evidence, that would not alter the value of graphic or constitutional formulæ. So far as these formulæ are concerned, it really makes little difference whether the atoms are actually attached to one another, or whether they are held in position by their mutual attractions while they are revolving about a centre of stability. Indeed, some assume that the conditions—thermal, optical, or electrical—necessary for the formation of a stable system determine whether a given atom can form a stable system with 1, 2, 3, . . . other atoms; otherwise stated, the valency of an element is determined by the necessity for harmonizing the peculiar motions of the combining atoms to form a stable molecular system. When Dalton was asked why an atom of carbon could take up one or two atoms of oxygen, but not three or four, he replied:

The reason I should assign is that in the state of CO_2 there are two atoms of oxygen combined with one atom of carbon, and a third or fourth atom of oxygen, however it may be attracted by the carbon, cannot join it without repelling one or more atoms of oxygen already combined. The attraction of the carbon is able to restrain the mutual repulsion of two atoms of oxygen, but not of three or four.

S. Brown expressed the same idea in 1843; he said: "The conception can perhaps be made still more lucid by the counter statement in astronomy that a sun cannot be overloaded with planets." Brown's view of valency shows that it is not necessary to postulate a distinct force emanating from the atoms in order to explain how, say, HCl forms a stable molecular system, while HCl_2 and H_2Cl do not form stable molecular systems. If such systems were momentarily formed, the supernumerary atoms would be immediately flung off. After trying motions and unions of every kind, the atoms fall into those favourable arrangements which can persist.

as stable molecules. There may, of course, be a number of different stable systems corresponding with the different stable molecules of, say, iron and chlorine, FeCl_2 and FeCl_3 . The plausibility of this hypothesis, of course, is not a proof that it is true. Other hypotheses are indicated in the last chapter of this book.

5. The energy of the atoms.—Each elementary atom, as we have seen, presumably has its own definite charge of energy. This energy probably exists in the form of atomic motions, so that when one atom “unites” with another atom, each atom possibly gives up a part of its energy, or “absorbs” energy from some external source, so that the motions of the one atom may be co-mingled with the motions of the other atoms to form a stable molecular system.

6. What makes the atoms and molecules move? We do not know! How can matter of itself initiate motion, and particularly motion in a harmoniously working system? *Ignoramus*. In the words of C. Kingsley, “Everywhere, skin-deep below our boasted science, we are brought up short by mystery impalpable and by the adamantine walls of transcendental forces and incomprehensible laws.” Consequently, the kinetic theories of atoms, of molecules, of the planetary systems, and indeed of the solar system itself, all prescribe or postulate an initial state of motion which is self-sustained and self-regulated. Guesses at the birth-history of these motions have been whispered only by the poets. Thus, in the oft-quoted lines of Virgil:

Know first, the heaven, the earth, the main,
The moon's pale orb, the starry train,
Are nourished by a soul,
A bright intelligence, whose flame
Glows in each member of the frame,
And stirs the mighty whole.

Immortal Newton could get no further than this: “The motions which the planets now have could not spring from any natural cause.” It seems as if

Full many a secret in her sacred veil
Hath Nature folded. She vouchsafes to knowledge
Not every mystery, reserving much
For human veneration, not research.—ANON. (1551).

§ 12. The Effect of Molecular Attraction on an Expanding Gas —The Joule-Thomson Effect.

I here use the word “attraction” in a general way for any endeavour of what kind soever made by bodies to approach to each other; whether that endeavour arise from the action of the bodies themselves, or whether it arise from the action of the æther or of any medium whatsoever, whether corporeal or incorporeal, anyhow impelling bodies placed therein towards one another.—ISAAC NEWTON.

W. Cullen (1755) seems to have been the first to notice that the temperature of air is decreased by rarefaction, and increased by compression; and J. Dalton attempted to measure the effects. In a general way, it may be said that if a gas, whose molecules exert no attraction on one another, be confined in a suitable vessel, and compressed, the mechanical work employed in compressing the gas is equivalent to the product of the

pressure into the change in volume. This energy is transformed into an equivalent amount of heat which raises the temperature of the gas. On the other hand, if the gas expands against atmospheric pressure, the gas will be cooled because the gas itself has done a certain amount of work equivalent to the product of the atmospheric pressure into the change in volume.

No heat is developed when an ideal gas expands into a vacuum because no external work is done by the gas. This was established experimentally by some early experiments by J. L. Gay-Lussac (1807), and by J. P. Joule (1845). Compressed air was allowed to expand into an evacuated vessel, and the result, as Joule expressed it, was as follows: "No change of temperature occurs when air is allowed to expand in such a way as not to develop mechanical power." Hence, it was also inferred that no work is performed under these conditions against inter-molecular attractions.

Our study of Boyle's and Charles' laws has taught us that inter-molecular attractions occur with most gases. Hence, this latter deduction might be questioned. The experiments, however, were not sufficiently sensitive to detect the small change of temperature which occurs when a gas expands *in vacuo*, so that although no external work is done by the gas, internal work is done against molecular attraction. The molecules are torn apart, so to speak, against the (feeble) attractive force drawing them together. This involves an expenditure of energy—work must be done.

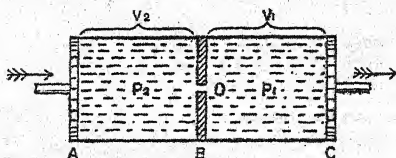


Fig. 41.—Joule-Thomson Effect.

Later, in a more delicate experiment, J. P. Joule and W. Thomson (Lord Kelvin)—1852-62—forced a steady stream of gas under a pressure p_2 slowly along a tube, A, Fig. 41, in the direction of the arrows, through a small orifice, O, where it expanded against the pressure p_1 . For the sake of simplicity, suppose the tube AB has unit sectional area, and that it is made of some material which does not conduct heat away from the gas. Two phenomena occur: (1) the gas is slightly heated by friction as it passes through the orifice O; and (2) the gas is cooled as it passes through O against a pressure p_1 .

Suppose a piston A, Fig. 41, moves from left to right so as to drive a volume of air v_2 at a pressure p_2 into the compartment BC. The work done on the gas is obviously $p_2 v_2$. Similarly, the work done by the gas as it pushes the piston from, say, B to C through a distance v_1 will be $p_1 v_1$. Hence, if the gas obeys Boyle's law, we shall have $p_1 v_1 = p_2 v_2$, and there will be no variation of temperature of the gas on the side AB and BC. If, however, work be done against molecular attraction during the expansion of the gas, the work of expansion on the side BC will exceed the work of compression on the side AB. The extra work will absorb heat from the gas itself. Hence, the gas on the side BC will be cooled below the temperature of the gas on the side AB. In Joule and Thomson's experiments, the temperature of carbon dioxide, nitrogen, oxygen, and air fell about 1° , while the temperature of hydrogen gas rose about 0.039° above the tem-

perature of the gas on the side AB . If, however, the experiment be conducted at a lower temperature, hydrogen gas behaves like the other gases, and is cooled. The change of temperature which occurs when a gas is driven through a small orifice is called the Joule-Thomson effect. The theoretical fall of temperature when the pressure falls from p_2 to p_1 is about $\frac{1}{4}^\circ$ per atm. difference $p_2 - p_1$ or more exactly, if T be the absolute temperature,

$$\text{Fall of temperature} = 75.35(p_2 - p_1)/T.$$

If carbon dioxide at 4 atm. pressure at 0° in passing through a porous plug suffers a fall of pressure to one atm., the fall of temperature will be $75.35 \times 3 \div 273$, or 0.828° .

§ 13. The Liquefaction of Gases.

When a thing is possible according to theory, and only practical difficulties oppose its realization, it is infinitely probable that these difficulties are not insurmountable.—G. CLAUDE.

When a gas, passing through an orifice O , Fig. 41, and cooled by the Joule-Thomson effect, is made to circulate around the tube leading to the compressed gas to the orifice, the gas issuing from the orifice will be cooled still more. The "self-intensive" or cumulative systems for cooling gases, elaborated by C. Linde, W. Hampson, and C. E. Tripler, between 1894-95, are based upon this principle. The idea will be understood after an examination of Fig. 42. The air to be liquefied—freed from carbon dioxide moisture, organic matter, etc.—enters the inner tube of concentric or annular pipes, A , under a pressure of about 200 atmospheres. This tube is hundreds of yards long and coiled spirally to economize space. By regulating the valve C the compressed air suddenly expands in the chamber D . The air thus chilled passes back through the tube B which surrounds the tube A conveying the incoming air. The latter is thus cooled still more. The gas passes along to the pumps where it is returned with more air to the inner tube. In this manner, the incoming air at 200 atmospheres pressure is cooled more and more as it issues from the jet O . Finally, when the temperature is reduced low enough, drops of liquid air issue from the jet. The tubes must all be packed in a non-conducting medium—wool, feathers, etc.—to protect them from the external heat.

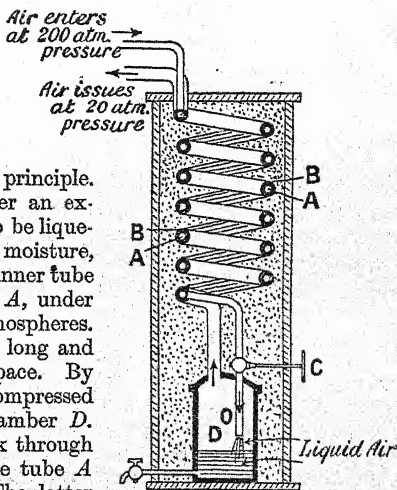


FIG. 42.—Linde's Apparatus for Liquefying Air (Diagrammatic).

Preserving liquid air.—There is a far greater difference between the

temperature of liquid air (about -190°) and ordinary atmospheric air, than between the temperature of ice and boiling water. The preservation of liquid air is thus a far more difficult problem than would be involved in preventing cold water boiling away while surrounded by a steam jacket at 200° . James Dewar solved the problem by keeping the liquid air in a double (or triple) walled vessel with the space between the walls evacuated, Fig. 43. Glass is a poor conductor, and a vacuum is a non-conductor. Hence, the liquid in the inner vessel can receive heat only from above, and by radiation. The glass walls of the evacuated space are also silvered to reduce the effects of radiant heat. Still air is a very bad conductor, so that the open end of the vessel is plugged lightly with cotton wool in order to reduce the ingress of heat from outside to a minimum. In this way, liquid air is transported by rail, etc., with a surprisingly little loss.

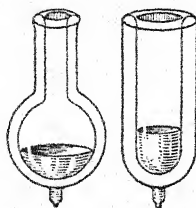


FIG. 43.—Dewar's Flasks.

the open end of the vessel is plugged lightly with cotton wool in order to reduce the ingress of heat from outside to a minimum. In this way, liquid air is transported by rail, etc., with a surprisingly little loss.

§ 14. The Manufacture of Oxygen and Nitrogen from Liquid Air.

Linde's process.—In Linde's process (1895), purified air is compressed to about 200 atmospheres, and driven along a pipe which divides at *A*, Fig. 44, into two streams and then passes down the interior tubes of a double set of annular or concentric pipes similar to the worm tube, Fig. 42. The two inner tubes finally unite into one single pipe, *B*. The air then passes through a spiral *S*, via the regulating valve *R*, and finally streams at *C* into the collecting vessel. The action is here similar to that described in the process for the liquefaction of air, Fig. 42. After a time, the air is liquefied in the collecting vessel, about the spiral *S*. The more volatile nitrogen boils off more rapidly than the oxygen. Hence, a gas rich in nitrogen passes up one of the two annular outer pipes as indicated on the left of Fig. 44. The liquid rich in oxygen is kept at a constant level by means of the valve, and thus the rate at which the liquid air in the collecting vessel is allowed to boil is also regulated. The oxygen passes from this tube on the right of Fig. 44 along the outer annular pipe, and finally emerges from the apparatus whence it is pumped into cylinders, etc., for use. If the valves are all properly regulated, the inrushing air is cooled by the counter currents of oxygen and nitrogen. The two latter gases pass along the tubes as indicated in the diagram. The tubes, etc., are all well insulated with non-conducting materials—feathers, wool, etc. By this process oxygen can be obtained as pure as is commercially desired,

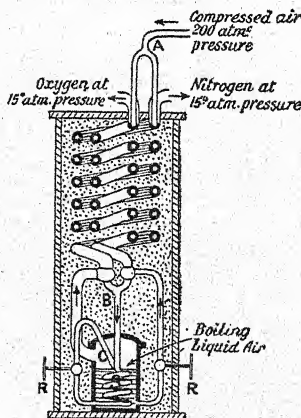


FIG. 44.—Linde's Apparatus for the manufacture of Oxygen from Liquid Air (Diagrammatic)

but the escaping nitrogen contains over 7 per cent. of oxygen. The original apparatus has been much improved, but Fig. 44 illustrates the underlying principle very well. As G. Claude has said: "We must salute in this apparatus the archetype of all the later machines, for it was the first to demonstrate that the manufacture of oxygen from liquid air is commercially possible."

Claude's counter-current process of rectification.—An improvement on Linde's process, by G. Claude (1903), enables practically pure oxygen and nitrogen to be obtained. When liquid air evaporates, the nitrogen—more volatile than the oxygen—escapes first, and the gas which comes from the liquid during the earlier stages of the evaporation contains so little oxygen that it will extinguish a lighted taper; as evaporation continues, the escaping gas becomes richer and richer in oxygen until at last the escaping oxygen is almost free from nitrogen. Similarly, if oxygen be bubbled through liquid air, the escaping gas contains about 93 per cent. of nitrogen—oxygen condenses from the rising bubbles, and nitrogen takes its place.

A diagrammatic sketch of Claude's apparatus is shown in Fig. 45. The cooled and purified air enters the lower part of the apparatus at a pressure of about 5 atm. and rises through a series of vertical pipes *P* surrounded by liquid oxygen, where it is partially liquefied. The liquid containing about 47 per cent. oxygen and 53 per cent. of nitrogen drains into the lower vessel *A*. The vapour which has survived condensation enters *B* and then descends through a ring of pipes *C* arranged concentrically about the set previously described. Here all is liquefied. The liquid which ultimately collects in this vessel *D* is very rich in nitrogen.

The pressure of the vapour in the central receptacle forces the liquid nitrogen to enter the summit of the rectifying column *E*, and the liquid, containing 47 per cent. of oxygen, is likewise forced to enter the rectifying column at *F* lower down. The pressures and rates of flow are regulated by the cocks *RR'*. The liquid nitrogen is 3° or 4° lower in temperature than the liquid rich in oxygen. Nitrogen evaporates from the down-streaming liquid, and oxygen condenses from the up-streaming gases. The heat supplied by the condensation of oxygen helps on the evaporation of nitrogen. Consequently, the descending liquid gets progressively richer and richer in oxygen, and the ascending gases richer in nitrogen. The liquid oxygen drains into the receptacle *G*, and is there evaporated by the latent heat of the gases condensing in the tubes. Finally, oxygen containing from 2 to 4 per cent. of nitrogen passes from the oxygen exit, and nitrogen containing 0.2 to 1 per cent. of oxygen escapes at the top of the rectifying column.

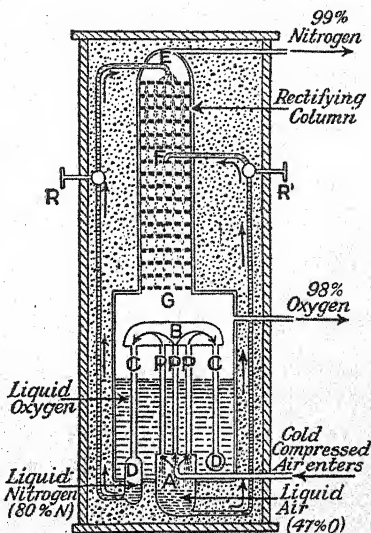


FIG. 45.—Claude's Apparatus for the separation of Oxygen from Liquid Air (Diagrammatic).

These processes enable oxygen and nitrogen to be made comparatively cheaply—one ton of coal for driving the compression apparatus is said to

furnish one ton of oxygen and four tons of nitrogen. As Linde says, "the heat is eliminated from the air exclusively by the expenditure of internal work," that is, the work required for the liquefaction of these gases is solely spent in separating the molecules of the gases from one another against their inter-molecular attractions.

Questions.

1. On the basis of the Kinetic Theory deduce a relationship between the pressure of a gas and the velocity and density of its molecules. Calculate the molecular weight of an unknown gas which under precisely similar conditions takes 1.117 times as long as oxygen to diffuse through an aperture.—*Sheffield Univ.*

2. How do you understand that the principle of the conservation of energy applies to a case on which a coiled spring (as of brass) is compressed within a vessel of glass or porcelain, and then completely dissolved in acid?—*Massachusetts Inst. Technology.* Hint: more heat is generated when a coiled spring is dissolved than is the case with an uncoiled spring.

3. Describe, with an explanation of the theoretical principles involved, the process for obtaining oxygen on an industrial scale by liquefying atmospheric air.—*Board of Educ.*

4. E. Molinari has said: "Newton believed that the pressure produced by compressed gases on the walls of the retaining vessel was due to the action of repulsive forces exercised by the gaseous particles on one another. This hypothesis is untenable because, if the gaseous particles were endowed with this repulsive force, then, in expanding freely without doing any work, this force should be transformed into heat, and thus, on expanding a gas, should be warmed." A commentator disagrees with this, adding: "Actually the work done would absorb heat from the gas itself and make it cooler." What do you think?

5. What is meant by a chemical change? In what respects does chemical action differ from gravitation? What properties of matter do we define as chemical properties, and what are physical properties?—*Punjab Univ.*

6. What hypothesis has been formed as to the physical nature of gases, and how does it explain the expansion of a gas by heat and the relation between the density, the molecular mass, and the diffusion rate of a gas.—*New Zealand Univ.*

7. What do you understand by the kinetic theory of gases? how can you utilize your knowledge to formulate a proof of Avogadro's hypothesis?—*Punjab Univ.*

8. How is liquid air prepared in quantity? Diagrammatically sketch the apparatus employed, and explain fully the principles involved.—*Punjab Univ.*

9. Explain the characteristic differences between solids, liquids, and gases in terms of the kinetic theory of molecules.

10. What is our present opinion concerning the physical nature of a gas? Point out how the law of Boyle, expressing the relation of the volume of a gas to the pressure, of Dalton (or Charles) connecting the volume of a gas with temperature, and of Graham connecting the rate of movement of a gas with its density, receive their explanation by means of this conception.—*New Zealand.*

11. What is the thermal evidence that the attraction between the molecules of the ordinary gases under standard conditions is small?—*Science and Art Dept.*

CHAPTER VIII

OXYGEN

Atomic weight, $O = 16$; molecular weight, $O_2 = 32$; bi- or quadri-valent. Melting point, -22° ; boiling point, -182.5° ; critical temperature -119° . Relative vapour density ($H_2 = 2$), 31.762 ; (air $= 1$) 1.1045 . One litre of oxygen at 0° , 760 mm., and at latitude 45° at sea-level, weighs 1.4292 grams, and 1 gram occupies 0.6997 litre.

§ 1. Oxygen—Occurrence and History.

Occurrence.—About one-fourth of the atmospheric air, by weight, consists of free oxygen, and water contains nearly 89 per cent. of combined oxygen. Oxygen also forms a material part of rocks. It is estimated that nearly one-half of the total weight of the rocks which make up the half-mile crust of the earth is oxygen, see p. 22.

History.—Carl Wilhelm Scheele's laboratory notes, preserved in the Royal Academy of Science, at Stockholm, are said to prove that Scheele discovered oxygen gas some time before 1773. Scheele called the gas "fire-air" and "vital air." Scheele made oxygen by heating red oxide of mercury; sulphuric acid and manganese dioxide; nitre; and some other substances. Scheele did not publish an account of his work until 1777. Meanwhile Joseph Priestley independently prepared the same gas, which he called "dephlogisticated air," while examining the effect of heat upon a great variety of substances confined in a cylinder (A, Fig. 46) along with mercury, and inverted in a trough of mercury, somewhat as in Fig. 46. Priestley focused the sun's rays upon the different substances by means of a "burning lens of 12 inches diameter, and 20 inches focal distance." Priestley announced his discovery of oxygen in these words: "On the 1st of August,



Fig. 46.—Priestley's Experiment.

1774, I endeavoured to extract air from *mercurius calcinatus per se*; ¹ and I presently found that by means of this lens, air was expelled from it very readily. Having got about three or four times as much as the bulk of my materials, I admitted water to it and found that it was

¹ That is, mercuric oxide, or red oxide of mercury.

not imbibed by it. But what surprised me more than I can well express was that a candle burned in this air with a remarkable brilliant flame."

Many erring steps have stumbled on the threshold of the discovery of oxygen; for instance, Eck de Sultzbach, in 1489, knew that red oxide of mercury gave off a "spirit" when heated; had he named and isolated the "spirit" he would have been credited with the discovery of oxygen. In 1678, O. Borch prepared the gas by heating saltpetre; in 1727, S. Hales collected the gas obtained by heating saltpetre; P. Payen obtained it in 1774; and J. Priestley himself obtained it the same year. These are not usually considered to have been discoveries of the gas because no attempt was made to determine the specific properties of the product. There are also indications in old books that the Greeks knew about oxygen in the fourth century; and that the Chinese were acquainted with the gas long before Priestley's and Scheele's experiments.

§ 2. The Preparation of Oxygen.

Oxygen can be made by simply heating mercuric oxide. Many other oxides are available—*e.g.* gold and silver oxides decompose at lower temperatures than mercury oxide, while manganese peroxide (pyrolusite) decomposes at a higher temperature. Unlike silver, mercury, and gold oxides, manganese oxide does not break down into the corresponding metal and oxygen, but rather into a complex oxide similar in composition to the mineral hausmannite— Mn_3O_4 . The reaction is symbolized: $3MnO_2 = Mn_3O_4 + O_2$. Gold, silver, and mercuric oxides are not often used on account of the expense. It might be added that the red oxide of mercury decomposes more rapidly if it be mixed with powdered platinum, ferric oxide, cadmium oxide, etc.

Mercuric oxide process.—When a supply of oxygen is required for experimental purposes, the apparatus can be arranged differently from

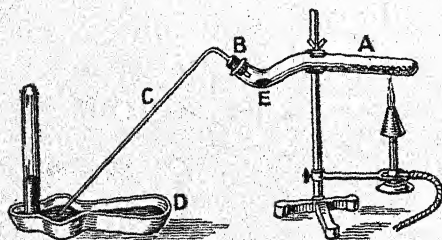


FIG. 47.—The Preparation of Oxygen.

Priestley's plan. The mercuric oxide can be placed in a hard glass tube A, Fig. 47, bent as shown in the diagram, and fitted with a cork B, and bent-glass delivery tube C. One end of the delivery tube dips in water, Fig. 10, or in special cases, in mercury, Fig. 47. The mercury gas-trough shown at D

can be worked with 5 lbs. of mercury, and it is convenient for collecting small quantities of gas when it is desirable to keep the gas out of contact with water. In most cases, of course, the ordinary water pneumatic trough, Fig. 10. will be used. The temperature of the vessel containing the mercuric oxide is gradually raised. The air in the tube is first driven off and allowed to escape. Later, globules of mercury begin to collect in the upper part of the tube, and oxygen is given off. The

mercuric oxide decomposes: $2\text{HgO} = 2\text{Hg} + \text{O}_2$. The oxygen is collected for examination; the mercury collects in the bend *E*, and there is no risk of mercury running back on to the hot glass and cracking the tube. 10 grams of mercuric oxide give not quite half a litre of oxygen.

The action of heat on potassium chlorate.—Potassium chlorate is a white crystalline solid which melts to a clear liquid when heated to about 340° . At about 10° higher the melted chlorate appears to boil, because bubbles of oxygen gas are copiously evolved. The potassium chlorate is decomposing. When the bubbling ceases, the molten mass begins to "thicken" or solidify. The potassium chlorate has decomposed into potassium perchlorate, potassium chloride, and oxygen. If the temperature be raised still further—over 600° —the mass again melts to a clear liquid and the potassium perchlorate decomposes, giving off more oxygen. The final products of decomposition are potassium chloride and oxygen. Hence, potassium chlorate can be used in place of mercuric oxide for the preparation of oxygen gas. Ten grams of potassium chlorate will give nearly $2\frac{1}{2}$ litres of oxygen. The amount of oxygen given off by heating a definite weight of mercuric oxide or potassium chlorate can be determined by placing about a gram of, say, potassium chlorate, dried at 150° , in a hard glass tube *C*, Fig. 48. Care must be taken that no chlorate sticks to the side of the tube or it may escape the action of heat. The increase in weight of the tube before and after the introduction of the chlorate represents the amount of chlorate used. After the chlorate has been heated, very gently at first, and afterwards very strongly, cool, weigh the tube, and also measure the volume of oxygen in the burette. The manipulation with the burette is the same as described for Fig. 34, p. 115. The loss in weight of the tube represents the amount of oxygen which has been expelled; and this number will agree closely with the weight of the gas calculated from the volume measured in the gas burette, if the measurements are accurate.

The potassium chlorate decomposes completely at a much lower temperature if it be mixed with manganese dioxide. Manganese dioxide, when heated alone, does not give off oxygen below 400° ; potassium chlorate alone does not give off oxygen below about 340° ; a mixture of the two gives off oxygen at about 200° . After the action, manganese dioxide still remains, but the potassium chlorate has decomposed into

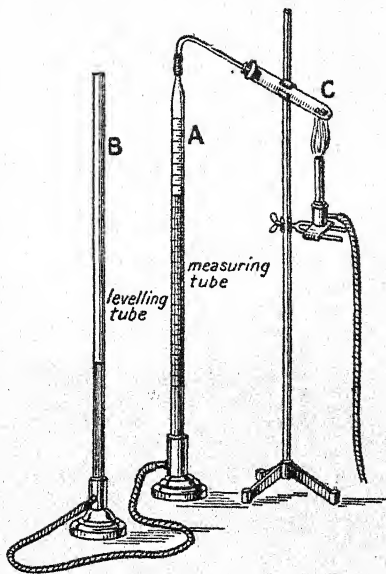


FIG. 48.—The Determination of Oxygen evolved by heating Potassium Chlorate.

potassium chloride and oxygen. Manganese dioxide can be recovered from the residue by lixiviating the mass with water. The water dissolves the potassium chloride, and leaves the manganese dioxide as a residue. Cobalt and nickel oxides, like manganese oxide, accelerate the decomposition of potassium chlorate.

If potassium chlorate be suddenly heated to a temperature above that at which decomposition occurs, the salt may detonate in an open vessel under ordinary pressure. Some disastrous explosions have been produced by potassium chlorate. M. Berthelot's experiment (1899) illustrates the explosive nature of this salt.

One end of a glass rod is drawn out into a thread, and the narrow end is dipped several times in melted potassium chlorate so that each layer of salt solidifies before the rod is dipped again. When a bend has been formed at the end of the rod, dip the rod into a test-tube heated red hot at one end so that the salt is about a centimetre from the bottom of the tube. Take care not to touch the sides of the tube. As the chlorate melts, it slowly drops to the bottom of the test-tube; each drop of chlorate as it falls explodes with a sharp detonation.

Potassium chlorate process for making oxygen.—For regular experimental work, oxygen is prepared by heating a mixture of potassium chlorate

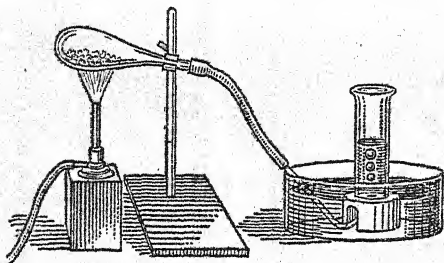


FIG. 49.—The Preparation of Oxygen.

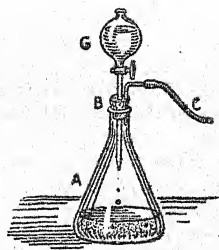


FIG. 50.—The Preparation of Oxygen.

(not powdered) with its own bulk of manganese dioxide ¹—oxygen mixture—in a wide-necked Florence flask, or a retort, or a special copper "oxygen flask," ² fitted with a wide delivery tube, because the gas is liable to come off rapidly in rushes. The flask is best clamped while tilted slightly downward towards the mouth, as indicated in Fig. 49, because a considerable amount of moisture is usually discharged from the mixture, and there is a risk of the moisture trickling back and cracking the glass. The gas is collected over water as in the case of hydrogen.

Peroxide method.—A third method of preparing the gas depends upon the fact that sodium dioxide is decomposed by water into sodium hydroxide and oxygen: $2\text{Na}_2\text{O}_2 + 2\text{H}_2\text{O} = 4\text{NaOH} + \text{O}_2$. Place about 10 grams of dry sodium dioxide in a dry 200 c.c. Erlenmeyer's flask, A, Fig. 50. The flask is also fitted with a stopper B bored with two holes. One hole is fitted with a delivery tube C, and the other hole with a tap

¹ If the manganese dioxide contains carbonaceous matters, an explosion may occur. Hence the manganese dioxide should be tested by heating a little with potassium chlorate before a large quantity is heated.

² 40 grams of the oxygen mixture may be used with a 250 c.c. flask.

funnel *G* containing water. The water from the funnel is allowed to fall, drop by drop, on to the peroxide. Each drop of water produces a definite amount of gas, so that the rate of evolution of the gas can be regulated by the rate at which water is allowed to drop from the funnel. No external application of heat is required. $6\frac{1}{2}$ grams of the peroxide give about a litre of oxygen. *Ozone* is a commercial preparation consisting of sodium dioxide mixed with a little catalytic agent like colloidal manganese dioxide, and is used for the making of oxygen on a small scale.

Permanganate process.—Many other methods are available for the preparation of oxygen. Heating sulphuric acid with manganese dioxide, chromic acid, potassium dichromate, potassium permanganate, or other salts rich in oxygen. A 10 per cent. solution of hydrogen dioxide and a concentrated solution of potassium permanganate gives off oxygen when acidified with sulphuric acid. It can also be made on a small scale by heating a mixture of, say, 20 grams of potassium permanganate with 80–100 c.c. of dilute sulphuric acid (one volume of the concentrated acid with four volumes of water) in a flask with a delivery tube and safety funnel as in, say, Fig. 103. Oxygen begins to come off when the temperature is about 50° , and continues in a steady stream. Ten grams of the permanganate with between 40–50 c.c. of the dilute sulphuric acid give just over a litre of gas.

Industrial preparation.—The particular process to be employed must be determined by cost and convenience. If but a few litres of gas, not specially purified, are required, cost is not very serious, and convenience is perhaps the most important factor; if pure oxygen be required, a complicated apparatus may be needed, and neither cost nor labour must be spared. An elaborate apparatus may be needed to remove traces of impurities—say, traces of ozone and chlorine from the oxygen. Pure potassium chlorate alone will give a gas of a high degree of purity. This method of preparation was used by Morley—p. 62—in his work on the atomic weight of oxygen. If large quantities of gas are needed, say for industrial purposes, the cost factor is of prime importance. Generally speaking, the success of industrial operations depends upon the ability of the chemist to manufacture his products cheaply. Oxygen was formerly made on a manufacturing scale by Brin's process (1881). This depends upon a very interesting reaction. When barium oxide— BaO —is heated in air to about 500° , it is rapidly oxidized to barium dioxide: $2\text{BaO} + \text{O}_2 = 2\text{BaO}_2$. If the barium dioxide be heated to a still higher temperature, 800° , the oxygen is given off and barium oxide remains as a residue: $2\text{BaO}_2 = 2\text{BaO} + \text{O}_2$. The barium oxide can be reoxidized and used over and over again, provided the air be freed from carbon dioxide, organic matter, dust, and any substance which forms a compound with barium oxide which is not decomposed under the given conditions. The regulation of the temperature offered practical difficulties which were overcome by keeping the temperature constant in the vicinity of 700° . Barium oxide is then transformed into the dioxide if the pressure of the atmosphere be about 2 kilos per square cm.¹ The peroxide is decomposed into the oxide and oxygen at the same temperature under a reduced pressure—about 0.05 kilogram per square cm. The gas pumped off under these conditions contained about 90–96 per cent. of oxygen, and 4–10 per cent. of nitrogen.

¹ Normally the atmospheric pressure is 1.033 kilograms per square cm.

Brin's process could not compete successfully against the cheaper method of preparation by the fractional distillation of liquid air (p. 152). Nearly all the oxygen on the market is now obtained by the liquid-air process. Very little is obtained by the electrolytic process indicated on p. 114. The oxygen is pumped into steel cylinders under a pressure of 100–150 atmospheres, and sold as "compressed oxygen." The gas may be obtained from the cylinders at any desired rate by opening the valve.

§ 3. Catalysis.

Materializing abstractions is a vice of thought.—H. S. R. ELLIOT.

The action of manganese dioxide on the decomposing potassium chlorate is very curious. It acts as a *stimulant*. We do not know precisely how the manganese peroxide does its work, although we can form a rough idea of what is taking place. Many other oxides act similarly, but not quite so vigorously—*e.g.* ferric, copper, cobalt, or nickel oxide may be used in place of manganese oxide. It is quite a common thing to find that the speed of reactions is accelerated or retarded by the presence of a "foreign" substance whose composition at the end of the reaction is the same as it was at the beginning. We met two cases in connection with the combustion of hydrogen—platinized asbestos, and moisture. These agents are conveniently grouped together as "catalytic agents," and the general phenomenon is called "catalysis." It must be clearly understood that "catalytic reactions" is simply a term for grouping those reactions whose speed is modified, or for those reactions which can be started by the presence of a small amount of a substance which is found to possess, at the end of the reaction, the same chemical composition as it had at the beginning. The catalytic agent may be affected by interaction with the products of the reaction, etc.—see "Consecutive reactions" (§ 5). The word "catalysis" itself explains nothing. To think otherwise would lay us open to Mephistopheles' gibe:

A pompous word will stand you instead
For that which will not go into the head.

This means that too much trust must not be placed in words. It is just when ideas fail that a word comes in most opportunely. There is no difficulty in covering an obscure idea by a word so that the word appears to explain the idea. In passing back from the word to the idea, it becomes easy to believe that the "subjective abstraction has an objective existence," or that because there is a word, something real must lie behind the word.

These remarks about the term "catalysis" might be applied, *mutatis mutandis*, to many of the terms in common use in chemistry—"passive resistance," "chemical affinity," the "ions" of the ionic theory, etc. Ostwald ingeniously compares the action of a catalytic agent with the action of oil on a machine, or of a whip on a sluggish horse. Ostwald, and his followers, believe that the reaction *must* be actually in progress before the catalytic agent can act. This limitation is quite arbitrary and, so far as we can see, does not agree with all the facts. Catalytic agents can start, accelerate, or retard the speed of chemical reactions. Some have tried to evade the difficulty by calling those reactions whose speed is merely accelerated by the catalyst, *catalytic reactions*; and those reactions

which cannot start unless their chemical energy is released or loosened in some way, trigger reactions—*vide* period of induction.

§ 4. The Properties of Oxygen.

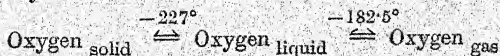
Oxygen, like air, is, at ordinary temperatures, a colourless, tasteless, and odourless gas. It is a little heavier than air :

Litre of normal air (760 mm. and 0°)	1.292 grams.
Litre of oxygen (760 mm. and 0°)	1.429 grams.

Oxygen is appreciably soluble in water—100 volumes of water, at 0°, can dissolve nearly 5 volumes of oxygen under a normal pressure of 760 mm., and at 20° about 3 volumes of the gas are dissolved. Fish are dependent upon the air dissolved in water for the oxygen they need for respiration. Animals are dependent upon the oxygen in air for respiration. Air normally contains 21 per cent. of oxygen, and if the proportion falls to 17 per cent. no known deleterious effects have been observed. Indians living in the upper reaches of the Andes breathe air which has oxygen equivalent to air with 12 per cent. of oxygen at ordinary atmospheric pressure. When the proportion of oxygen falls below this value the air is getting dangerous, and although air with less than 7 per cent. of oxygen can be breathed with impunity for a short time, life will soon be extinguished. A candle is extinguished in air with less than 17 per cent. of oxygen. A mouse soon dies if placed in an atmosphere deprived of oxygen. Pure oxygen can be breathed for a short time without harm, and oxygen is used medicinally for the resuscitation of persons who have been suffocated, or suffering from carbon monoxide poisoning, etc., where, owing to the enfeebled action of the lungs, the blood is not sufficiently aerated; and where it is necessary to stimulate a person with an internal injury—*e.g.* a broken rib—which would prevent artificial respiration. The prolonged inhalation of oxygen soon raises the temperature of the body dangerously high. An animal placed in ordinary or in compressed oxygen soon dies.

Metallic silver, gold, platinum, and palladium absorb oxygen at about 500°. Molten silver dissolves about ten times its volume of oxygen, and gives it up again on cooling. In cooling, a solid crust forms on the exterior surface; as the interior cools the gas bursts through the solid crust driving out a spurt of the still fluid metal—the phenomenon is called the “spitting” of silver. Molten platinum behaves in a somewhat similar way.

Oxygen has been condensed to a bluish-coloured mobile liquid at -119° under a pressure of 50 atmospheres. If the temperature exceeds -119° , oxygen cannot be liquefied by any pressure however great. The smallest pressure which will liquefy oxygen at -119° is 50 atmospheres. The temperature -119° is therefore the critical temperature, and 50 atmospheres is the critical pressure of the gas. Oxygen gas is feebly magnetic. Liquid oxygen boils at -182.5° , pressure 760 mm. The liquid has a specific gravity of 1.13, that is, liquid oxygen is 1.13 times as heavy as an equal bulk of water. Liquid oxygen can be frozen to a pale bluish-white solid not unlike snow in appearance. The solid melts at -227° , and has a specific gravity 1.43. With the notation used before :



Liquid oxygen is strongly attracted by a magnet, so that if a little liquid oxygen is placed in a cup beneath the poles of an electromagnet, it leaps up to the poles and remains permanently attached until all is evaporated. Mixtures of liquid oxygen and petroleum are violently explosive. It is said that a lighted candle falling into a bucket of liquid oxygen sent G. Claude to the hospital "in a very pitiable condition." Lamp-black soaked in liquid oxygen burns slowly when ignited, but explodes violently when detonated by a fulminate primer. Liquid air, or rather liquid air enriched with oxygen, furnishes the explosive *oxyliquite* when mixed with charcoal; 3 cm. cartridges charged with one part carbon, one part petroleum, and eight parts of liquid oxygen were tried experimentally when cutting the Simplon tunnel. The cartridges are exploded by an electric fuse and a mercury fulminate cap. The chief objection is that the cartridges must be used within three minutes after charging, or the oxygen will evaporate. This objection might be an advantage under some circumstances, since a mis-fired shot is harmless in a short time.

Oxygen is used in conjunction with hydrogen for the oxy-hydrogen flame (*q.v.*), and with acetylene for the oxy-acetylene flame used in welding. Thick steel plates can be cut with ease by playing a fine stream of oxygen on the red-hot metal. Advantages are claimed for enriching with oxygen the air-blast of a blast furnace, etc. Oxygen is used in bleaching, in the oxidation and thickening of oils to be used in making varnishes, linoleum, etc.

The great chemical activity of oxygen is well typified by Priestley's quaint observation, indicated on p. 155. A glowing splint of wood ("cedar splints") when plunged into oxygen bursts into flame, the carbon of the wood being oxidized to carbon dioxide (CO_2). The inflammation of a glowing splint is often used as a test for oxygen. A mixture of air and oxygen with less than 28–29 per cent. of oxygen does not re-ignite a glowing splint, and if the air has less than 16 per cent. of oxygen the glowing splint will be extinguished. Oxygen alone has no action on clear lime water, but after a splint has burnt in the gas, the clear lime water becomes turbid. Oxygen combines directly with most other elements, particularly at elevated temperatures, forming oxides. Iodine, bromine, fluorine, gold, platinum, and argon and its companions do not combine directly with oxygen; but oxygen combines indirectly with all elements excepting the argon group of elements, fluorine, and possibly bromine. If the metals be arranged in the order of their avidity or readiness to combine with oxygen, caesium, potassium, and sodium will be found at one end of the series, while platinum and the argon family will be found at the other end.

The direct combination of oxygen with some of the elements can be illustrated by placing small dry pieces of carbon, sulphur, phosphorus in deflagrating spoons, heating them until combustion begins, and then plunging each into a jar of oxygen.¹ The glowing piece of charcoal burns very brightly and forms a gaseous oxide—carbon dioxide, CO_2 . Sulphur burns with a lavender-blue flame, forming gaseous sulphur dioxide— SO_2 —which has the peculiar odour characteristic of burning sulphur. The reaction is symbolized: $\text{S} + \text{O}_2 = \text{SO}_2$; sulphur dioxide is soluble in water forming sulphurous acid— H_2SO_3 —which reddens blue litmus

¹ Quick-sealing fruit jars fitted with a deflagrating spoon make excellent gas jars for these experiments.

solution— $\text{H}_2\text{O} + \text{SO}_2 = \text{H}_2\text{SO}_3$. The experiment can be conducted as recommended by F. Rüdorff, Fig. 51. The two-necked globe is corked at one end and filled with oxygen at the pneumatic trough in the ordinary manner; one neck is allowed to dip in a beaker of water as shown in the diagram. When the deflagrating spoon, fitted with a tightly fitting cork, is inserted with the burning sulphur into the globe, the water, coloured with blue litmus, rises in the vessel as the sulphur dioxide is absorbed. At the same time the litmus changes in colour, showing the formation of an acid during the burning of the sulphur. Phosphorus burns in oxygen vigorously and brilliantly, forming a white cloud of phosphorus pentoxide— P_2O_5 . The reaction is represented: $4\text{P} + 5\text{O}_2 = 2\text{P}_2\text{O}_5$. The phosphorus pentoxide dissolves in cold water, forming metaphosphoric acid— HPO_3 . The reaction is written: $\text{P}_2\text{O}_5 + \text{H}_2\text{O} = 2\text{HPO}_3$. These reactions will be studied in more detail when the elements in question receive individual treatment.

It might be added that dry sulphur, dry phosphorus, and dry carbon burn with great difficulty or not at all in dry oxygen. In fact, perfectly dried substances often appear to be chemically inert, whereas they react vigorously if a trace of moisture be present.

To show the combustion of iron in oxygen gas, tie a tuft of "steel wool" to the end of a stout iron wire by means of a piece of steel wire. Heat the end of the wool in a Bunsen's flame until incipient combustion begins, and quickly plunge it into a jar of oxygen on the bottom of which a layer of water, sand, or asbestos paper has been placed. The wool burns with dazzling scintillations, the product of the reaction—iron oxide—falls to the bottom of the jar in fused globules. When cold, the oxide of iron resembles a blacksmith's hammer scale. It is called black or magnetic oxide of iron— Fe_3O_4 . The reaction is usually written: $3\text{Fe} + 2\text{O}_2 = \text{Fe}_3\text{O}_4$. The subject of oxidation and combustion will be resumed in a later chapter.



FIG. 51.—Rüdorff's Gas Jar.

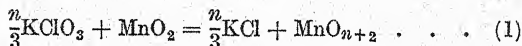
§ 5. Consecutive Reactions.

With the growth of knowledge, reaction represented by the older chemists by simple equations resolve themselves into reactions of greater and greater complexity.—ANON.

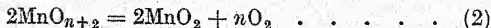
Let us return to the action of heat on potassium chlorate. The representation of a chemical reaction by means of an equation emphasizes the character of the initial and of the end products of the reaction, but it conveys no idea of the mechanism of the reaction—*how* the different materials interact to give the final products. The regular type of chemical equation shows but the beginning and the end of the reaction. There can be no doubt that quite a number of intermediate stages temporarily subsist before the drama of the reaction closes with the final act—the formation of the end products. There is plenty of evidence leading us to infer the

existence of a kaleidoscopic sequence of changing scenes during the progress of what are usually considered simple reactions. Some suppose that water has no more right to representation in the chemical equation than the glass of the vessel in which the reaction occurs. As we progress in our studies, we shall find that water profoundly modifies the properties of most substances with which it is in contact—*e.g.* the influence of moisture on the oxidation of hydrogen, sulphur, etc. "This," said C. L. Berthollet (1803), "is a striking illustration of the effect of words on the ideas we form and even on the results of observation. We begin by considering a solvent as the liquid employed in making solutions which we term 'mere mixtures of solvent and dissolved substances,' consequently, the attention is rarely directed to the influence of the solvent under other circumstances because, in them, it retains the name 'solvent.' It must, however, never be forgotten that all the substances present in a reacting system exercise an action; and if there are circumstances in which the solvent may be neglected, there are others in which it contributes efficaciously to the result."

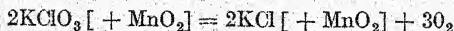
As a result of quite a number of experimental investigations on the decomposition of potassium chlorate, and a study of the available circumstantial evidence, we are able to get, in imagination, a peep behind the curtain which hides the course of the reaction. Firstly, it is not quite correct to say that the manganese dioxide is not changed in any way during the reaction because a microscopic examination of the manganese dioxide, before and after the reaction, shows that it has undergone a physical, if not a chemical, change—crystalline manganese dioxide apparently becomes amorphous. The manganese dioxide does appear to take part in the reaction in spite of the fact that it has the same chemical composition at the end as it had at the beginning. Secondly, the manganese dioxide is probably oxidized by the decomposing chlorate to form one of the unstable higher oxides of manganese, exactly what oxide we do not know. This uncertainty is expressed by writing the unknown oxide: MnO_{n+2} where the numerical value of n is not known with certainty. This stage of the reaction can then be represented by the equation:



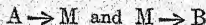
Thirdly, the unstable oxide produced by the oxidizing action of the potassium chlorate probably breaks down almost as soon as it is formed, regenerating the manganese dioxide, and liberating free oxygen:



The manganese dioxide so formed is again oxidized, and the oxide again decomposed regenerating manganese dioxide anew. This cycle of changes continues until the potassium chlorate is all decomposed. The opening and closing scenes are represented:



Equations (1) and (2), expressed in the most general form, indicate that we are dealing with a reaction in which

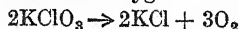


where A and B respectively denote the initial and final products of the reaction, and M the intermediate products. In the reaction just considered,

M is represented by MnO_{n+2} . Under the prevailing conditions, A does not form B directly. Consecutive reactions are those in which intermediate products are produced which do not necessarily appear as final products in the reaction. Consecutive reactions occur in stages; one stage must be in progress before another can start. The speed of formation of B from A obviously depends on the speed of the intermediate reactions. If the reaction $A \rightarrow M$ be very rapid, and $M \rightarrow B$ be very slow, the intermediate product M will accumulate in the system, and could be recognized and probably isolated. Several examples are known. On the other hand, if $A \rightarrow M$ be very slow, and $M \rightarrow B$ be very fast, it would be hopeless to look for intermediate products, and the evidence in support of the assertion that the reaction involves a sequence of consecutive or intermediate reactions must be circumstantial, not direct proof.¹

§ 6. Concurrent or Side Reactions.

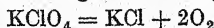
The fate of the potassium perchlorate.—Part of the chlorate decomposes into potassium chloride and oxygen:



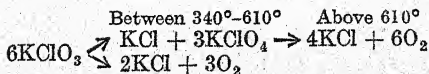
and part oxidizes another part of the chlorate into potassium perchlorate (KClO_4):



These two reactions proceed side by side—concurrently, yet independently. Measurements of the relative proportions of potassium perchlorate and oxygen formed at different temperatures show that the potassium perchlorate reaction proceeds nearly twice as fast as the other reaction. The lower the temperature, the greater the relative speed of the perchlorate reaction. Hence as the potassium perchlorate accumulates in the system, the molten mass becomes more and more viscid, and if the temperature be below the melting point of potassium perchlorate (610°), the mass solidifies when enough potassium perchlorate has accumulated in the system, even though the temperature be higher than the melting point of potassium chlorate (340°). When the temperature is raised high enough, the potassium perchlorate decomposes into potassium chloride and free oxygen. Here again the opening and closing scenes are represented by the equation:



But the whole reaction could perhaps be better represented by the scheme:²



¹ It will also be obvious that the same reasoning must apply in a longer series of intermediate reactions; $A \rightarrow M$; $M \rightarrow N$; $N \rightarrow B$, etc. Similarly, one or more of the intermediate reactions might be concurrent reactions (see below), or opposing reactions (p. 119).

² This view of the mechanism of the decomposition of potassium chlorate by heat shows how the relative proportions of potassium chloride, perchlorate, and oxygen depend on the temperature, and almost an infinite number of equations are possible. The students must bear this in mind when reading many text-books, for the reaction is often represented by complex equations. It can be shown that all so far proposed are special cases of the simple equations described in the text, and these are based on the work of W. H. Sodeau (1900-1903).

The final products of both reactions are potassium chloride and oxygen, and this is the sole justification for writing the equation $2\text{KClO}_3 = 2\text{KCl} + 3\text{O}_2$. When two or more reactions proceed simultaneously and independently side by side in the same system they are said to be concurrent or side reactions. If one of the reactions proceeds much faster than the others, it is said to be the **main reaction**; the others, **side reactions**. When potassium chlorate is heated with manganese dioxide, no potassium perchlorate has yet been detected among products of the reaction. Hence it is inferred that no potassium perchlorate is formed. If this be correct, the cyclic reactions between the manganese dioxide and potassium chlorate proceed rapidly at a temperature much lower than that at which the perchlorate reaction has acquired an appreciable velocity. In fine, the catalytic agent accelerates at least one of the two concurrent reactions.

It must not be supposed that the above outline gives a complete representation of this remarkable reaction. The products of the reaction may interact with themselves or with the catalytic reagent. In some cases part of the oxygen comes off as ozone, and the products of the reaction may contain a little chlorine. Traces of potassium permanganate have been detected among the residual products. The chlorine and potassium permanganate are probably formed by a reaction between the potassium chlorate and the manganese dioxide. As soon as the student gets beyond the kindergarten or pyrotechnical stages, chemistry becomes intellectually fascinating.

§ 7. The Origin of the Terms: Acid, Base, and Salt.

However convenient the classification of oxides into acids and bases might be for an elementary presentment of chemistry, a glance from the vantage ground of facts not usually referred to in elementary courses, shows such classifications to be imperfect and arbitrary to a degree.—
D. CARNEGIE.

The early chemists appear to have gradually learned to arrange certain substances into two groups according as these substances possessed certain qualities in common with vinegar or with wood ashes. The former were called **acids** (from the Latin, *acidus*, acid); and the latter, **alkalies** (from the Arabian, *alkali*, ashes of a plant), because the alkalies were generally obtained by calcining various materials and reducing them to ashes. Towards the end of the seventeenth century, Robert Boyle summarized the properties of acids as substances which (1) have a sour taste; (2) dissolve many substances (corrosive); (3) precipitate sulphur from alkaline solutions of sulphur; (4) change many vegetable blue colours (*e.g.* blue litmus) red; and (5) lose their acid characteristics when brought into contact with the alkalies. The alkalies were considered to be substances which (1) possessed detergent and soapy properties; (2) dissolved oils and sulphur; (3) restored vegetable colours reddened by acids; and (4) had the power of reacting with acids to produce indifferent substances.

The properties of acids and alkalies were thus opposed to one another; for when mixed together, the one neutralized the other. Salts were considered to be products of the interaction of acids and alkalies. It was soon found that some substances with alkaline qualities did not melt nor change

when heated, and were almost insoluble in water—these substances were called *earths*. In 1744 F. G. Rouelle employed the word *base* to include the earths, alkalis, metallic oxides ("calces"), and all substances which produce salts by reacting with the acids.

It was soon recognized that many substances could not well be grouped with the acids and bases, although they possessed qualities characteristic of acids or bases. Thus aluminium ammonium sulphate—alum—forms a solution with water which has a sour taste, deprives sodium hydroxide of its alkaline qualities, and turns blue litmus red; copper sulphate reddens blue litmus; sodium carbonate and sodium borate turn red litmus blue, etc. Conversely, substances may be grouped as acids and bases, even though they have no action on litmus, e.g. silicic acid, H_2SiO_3 , has no action, on blue litmus, and yet it is an acid; similarly copper oxide, CuO , is a base without action on red litmus.

§ 8. Acids.

Acid is rather the name of a function than the name of a substance.—ANON.

In his study of the properties of oxygen, Lavoisier noticed that when certain elements were burnt in oxygen, the resulting oxide forms an acid with water—e.g. carbon, sulphur, and phosphorus. Hence Lavoisier concluded (1777) that "oxygen is an element common to all acids, and the presence of oxygen constitutes or produces their acidity." Lavoisier considered oxygen to be the essential constituent of all acids. The very name *oxygen*, given to this element, was derived from Greek words signifying "the generative principle of acids"— $\delta\acute{\epsilon}\xi\acute{\iota}\varsigma$ (oxus), sour, and $\gamma\epsilon\nu\acute{\alpha}\omega$ (gennao), I produce—because "one of the most general properties of this element is to form acids by combining with many different substances; hence the German term for oxygen *Sauerstoff*, or acidifying stuff. For a time, *le principe oxygene* was a fetish with the French chemists; but, with increasing knowledge, it was found that Lavoisier's oxygen theory of acids led to confusion and error, and it was gradually abandoned by chemists when it was recognized that:

1. *Some oxides form alkalis, not acids, with water.*—E.g. sodium, potassium, and calcium oxides. As Humphry Davy expressed it, "the principle of acidity of the French nomenclature might now likewise be called the principle of alkaliescence."

2. *Some acids do not contain oxygen.*—C. L. Berthollet showed, in 1787, that hydrocyanic (prussic) acid is a compound of carbon, nitrogen, and hydrogen, but contains no oxygen; and he also came to a similar conclusion with regard to hydro-sulphuric—hydrogen sulphide. But for some time Lavoisier's reputation had more weight than Berthollet's facts. In 1810–11 Humphry Davy proved that hydrochloric acid is a compound of hydrogen and chlorine, and that no oxygen could be detected in the compound. In 1813 Davy also proved that hydriodic acid contained hydrogen and iodine, but no oxygen. Hence, added Davy, "acidity is not connected with the presence of any one element."

As a result of Davy's work, the acids came to be classed as *hydracids*—acids containing no oxygen; and *oxyacids*—acids formed from acidic oxides. In 1815 Davy suggested the possibility that hydrogen, not

oxygen, gives the acid characters to the acids; but he did not rush to the other extreme and say that all hydrogen compounds are necessarily acids. There is no one property which we can use as an absolute criterion or decisive test of acidity. In a crude sort of way, it can be said that *acids usually have a sour taste, are usually corrosive, redden the blue colour of vegetable substances (e.g. litmus), and contain hydrogen, part or all of which can be replaced when the acid is treated with a metal, metallic oxide, hydride, or carbonate.* Acids are known with a sweet taste, and which are not corrosive. Alum, as indicated above, does not contain replaceable hydrogen, and it would not therefore be classed as an acid, although it is sour, corrosive, and colours blue litmus red. But we are far from a satisfactory definition of acids, although, as we shall see, we can make a fair definition in terms of the ionic hypothesis.¹ Sodium bisulphate has a sour taste, is corrosive, reddens blue litmus, and contains replaceable hydrogen, but it is not usually regarded as an acid because of its mode of formation. Naturally the student delights in clear, sharp-cut definitions, and teachers of science have many temptations to frame definitions and draw boundary lines which do not exist in nature. "Definitions," said John Hunter, "are the most accursed of all things on the face of the earth."

§ 9. Salts.

There is perhaps no inquiry which has thrown so much light on a multitude of combinations which the want of method had permitted to be confounded as Rouelle's observations on the characters of salts.—C. L. BERTHOLLET.

In modern chemistry the word *salt* is a descriptive term applied to a distinct family of substances and not to any particular individual. In the kitchen, "salt" is colloquially applied to one specific individual, sodium chloride. A salt is produced by replacing all or part of the hydrogen of an acid by a metal or basic radicle. For instance zinc displaces the hydrogen, of sulphuric acid: $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$ forming zinc sulphate as indicated on p. 114. Hence C. Gerhardt (1843) defined acids to be "salts of hydrogen":

SO ₄ —BIVALENT RADICLE.		Cl—UNIVALENT RADICLE.	
Hydrogen sulphate (sulphuric acid)	H ₂ SO ₄	Hydrogen chloride (hydrochloric acid)	HCl
Zinc sulphate	ZnSO ₄	Zinc chloride	ZnCl ₂
Sodium sulphate	Na ₂ SO ₄	Sodium chloride	NaCl

Salts of the binary acids (*i.e.* acids compounded of two elements² like hydrochloric acid, hydrofluoric acid, etc.) are usually named by dropping the prefix "hydro-" and changing the termination "-ic" into "-ide." Thus the acids just named furnish chlorides, fluorides, etc. To show what chlorides, etc., are in question, the name of the corresponding element (or elements) is introduced as an adjective. Thus we have sodium chloride, potassium chloride, calcium chloride, etc. The names of the elements are thus used adjectivally in the same sense that the words

¹ Although the definitions in terms of the ionic hypothesis are not very different from the definitions in the text, so far as practical applications are concerned.

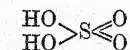
² Some radicles free from oxygen, *e.g.* CN or Cy, are regarded as if they were single elements. Why? Thus hydrocyanic acid—HCN—furnishes cyanides.

"stone," "brick," and "wood" prefixed to "house" are adjectival, and indicate the kind of house in question.¹

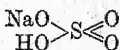
The salts of the ternary acids (*i.e.* acids with three elements) are named by changing the "-ic" termination of the acid "-ate," or the "-ous" termination of the acid into "-ite," and adding the word so obtained to the base or bases forming the salt. Thus sulphuric acid forms sulphates—*e.g.* sodium sulphate; nitric acid, nitrates—*e.g.* calcium nitrate; sulphurous acid, sulphites—*e.g.* ammonium sulphite; perchloric acid, perchlorates—*e.g.* potassium perchlorate; hypochlorous acid, hypochlorites—calcium hypochlorite; carbonic acid, carbonates—*e.g.* calcium carbonate, etc.

F. C. Rouelle, about 1754, distinguished the neutral salts of a given base from the acid salts with an excess of acid, and the question whether the proportions of acid and base in the two salts were constant so that there were but two combinations—one with a maximum and the other with a minimum quantity of acid—or if combinations could exist with intermediate proportions dependent on the circumstances under which the salts were formed was discussed by C. L. Berthollet and J. L. Proust in the controversy to which allusion has already been made.

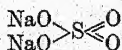
In normal salts all the displaceable hydrogen of the acid is replaced by the base. For instance, sodium sulphate— Na_2SO_4 —is a normal salt because all the replaceable hydrogen of sulphuric acid is displaced by sodium. In acid salts only part of the replaceable hydrogen has been displaced by a base, and the salt still contains replaceable hydrogen. For instance, acid sodium sulphate— NaHSO_4 —contains half the replaceable hydrogen of sulphuric acid, and half as many equivalents of sodium as normal sodium sulphate. If an acid contains two or more replaceable hydrogen atoms, it does not follow that all need be displaced by the same element. These ideas can be illustrated graphically—sulphur sexivalent:



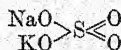
Sulphuric acid.



Acid sodium sulphate.



Normal sodium sulphate.



Sodium potassium sulphate.

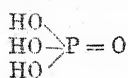
Sometimes the term "hydrogen" is used in place of "acid" for the acid salts, and sometimes the prefix "bi-" or "di-" is appended to the term for the acid in the salt. Thus "acid sodium sulphate" is also called "sodium hydrogen sulphate," "sodium bisulphate," as well as "mono-sodium sulphate," etc. The normal salts are sometimes called "neutral salts" in the sense that all the hydrogen has been "neutralized" or displaced from the acid. These salts, however, are not necessarily neutral to litmus—thus normal zinc or copper sulphates react towards litmus as if they were acids; borax, sodium nitrite, and normal sodium carbonate react as if they were alkalies.² Note the confusion in the use of the term

¹ Hence, some years ago, the name of the basic element used to be modified to give it an adjectival form; hydric chloride; potassic chloride; calcic chloride; etc. This system has been abandoned unless it is desired to distinguish between "-ous" and "-ic" compounds—*e.g.* ferrous chloride, and ferric chloride, etc.

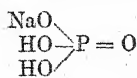
² Usually mercurous, mercuric, cupric, aluminium, chromic, ferric, stannous, stannic, antimonious, and bismuthous soluble salts with the common acids have an acid reaction—red den blue litmus; while the borates, carbonates, chromates, hypochlorites, nitrites, phosphates, silicates, sulphides, and sulphites of alkali metals have an alkaline reaction—turn red litmus blue.

"acid." The term may be used (i) as a noun, to denote a particular class of substances; (ii) as an adjective, to denote a certain quality or property—*e.g.* its effect on litmus, etc.; and (iii) as an adjective, to denote that a salt has replaceable hydrogen atoms. Many acid salts are *acid* to litmus, *e.g.* sodium hydrogen sulphate; others are *alkaline*, *e.g.* sodium hydrogen carbonate; others again are *neutral*, *e.g.* disodium hydrogen phosphate.

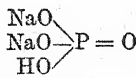
It is sometimes necessary to use the prefixes mono-, di-, tri-, . . . to discriminate between the different salts of one acid. Thus with phosphoric acid—phosphorus quinquivalent—



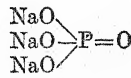
Phosphoric
acid.



Monosodium
phosphate.



Disodium
phosphate.



Normal or tri-sodium
phosphate.

It would be a mistake to assume that all the hydrogen of an acid is replaceable by a base. Thus, so far as we know, hypophosphorous acid— H_3PO_2 —has only one of its three hydrogen atoms replaceable by a metal. No one has ever prepared Na_2HPO_2 or Na_3PO_2 . The number of atoms of hydrogen in one molecule of an acid which are replaceable by a metal, or a radicle, is termed the basicity of the acid. Thus hydrochloric acid— HCl —is monobasic because each molecule of hydrochloric acid contains one replaceable hydrogen atom; sulphuric acid— H_2SO_4 —is dibasic; phosphoric acid— H_3PO_4 —is tribasic; and silicic acid— H_4SiO_4 —is tetrabasic. Hypophosphorous acid— H_3PO_2 —is monobasic.

§ 10. Neutralization.

A solution of sulphuric acid, like other acids, colours blue litmus red; and a solution of sodium hydroxide, like other alkalis, colours red litmus blue. It is possible to mix the acid with the alkali so as to furnish a solution which neither tastes nor reacts towards litmus like sulphuric acid nor like sodium hydroxide. If too much acid be present, litmus will be coloured red, and if too much alkali, blue. The mixture on evaporation furnishes a crystalline solid which neither colours blue litmus red, nor red litmus blue. The product of the reaction is said to be *neutral*, and the process of neutralization consists in adding an acid to an alkali, or an alkali to an acid, until a neutral substance is obtained. The result of the reaction is called a salt. The salt contains the metal of the alkali, and the radicle of the acid. The litmus used to determine the point of neutralization is called an indicator. Several other indicators besides litmus are available; *e.g.* phenolphthalein furnishes a pink coloration with alkaline solutions, and is colourless with acids and neutral solutions; methyl orange is yellow with alkalis; pink with acid and orange with neutral solutions. It will be observed that the determination of the neutral point is here referred *arbitrarily* to the behaviour of litmus because the mode of arrangement of substances into classes acidic, alkaline, and neutral substances depends on the indicator used. Salts of the heavy metals—alum, ferric chloride, and ferrous sulphate—are usually acid to both litmus and phenolphthalein, and neutral to methyl orange; borax, sodium or potassium hydrocarbonate, and sodium citrate are alkaline to

litmus and methyl orange, but neutral to phenolphthalein; while sodium phosphate is neutral to litmus and phenolphthalein, and alkaline to methyl orange. When phosphoric acid is titrated with sodium hydroxide it appears to be a monobasic acid if methyl orange be used as indicator; and as a dibasic acid with litmus or phenolphthalein as indicator. Again, boric acid does not affect methyl orange, but it changes both litmus and phenolphthalein; and an aqueous solution of potassium sulphite is neutral to phenolphthalein, and turns violet litmus blue.

Richter's work.—The above experiment coupled with numerous others with different acids and bases have proved that acids and alkalis unite to form salts in constant proportions. Many of the early chemists—G. Homberg (1699), A. L. Lavoisier (1782), H. Cavendish (1788), T. Bergman (1785), C. F. Wenzel (1777), etc.—had a more or less clear idea that a definite weight of a base neutralized a definite amount of a given acid: but J. B. Richter, in an important study of this subject between 1791 and 1802, demonstrated conclusively that the weights of the various acids which neutralize a certain fixed weight of one of the bases are the same for certain fixed weights of all the bases; and the same numbers hold good for the neutralization of the acids by the bases. This is Richter's law of proportionality. Consequently it is possible to assign equivalent numbers to the acids and bases. For instance, using modern data and terms:

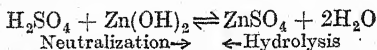
ACIDS.		BASES.	
	Equivalent weight.		Equivalent weight.
Hydrofluoric acid . . .	20·01	Ammonium hydroxide . . .	35·05
Hydrochloric acid . . .	36·47	Calcium hydroxide . . .	37·06
Sulphuric acid . . .	49·04	Sodium hydroxide . . .	40·01
Nitric acid . . .	63·02	Potassium hydroxide . . .	56·00

A table analogous with this, but with less accurate data, was calculated in 1802 by G. E. Fischer from J. B. Richter's data, and this was the first table of equivalent weights published. The weights of the acids in one column represent the amounts required to neutralize the quantity of any of the bases indicated in the other column: and conversely, the weights of the bases in the second column represent the amounts required to neutralize the quantity of any one of the acids indicated in the first column. Thus 56 grams of potassium hydroxide will neutralize 20·01 grams of hydrofluoric acid, 36·47 grams of hydrochloric acid, 49·04 grams of sulphuric acid, 63·02 grams of nitric acid, etc., and 63·02 grams of nitric acid will neutralize 35·05 grams of ammonium hydroxide, 37·06 grams of calcium hydroxide, etc. Richter mixed up much valuable work with several fantastic hypotheses, and he also "cooked" some of his results so that they represent what he thinks he ought to have obtained rather than what he actually observed. Such a procedure is quite antagonistic to the "spirit of science," and made chemists reasonably sceptical about the accuracy of the whole of Richter's work. It was thought, wrongly as it happens, *falsus in uno, falsus in omnibus* (false in one, false in all). Consequently Richter's generalization did not attract the attention it deserved.

It follows as a corollary from Richter's law that when two neutral salt solutions mutually decompose one another, the newly formed products are also neutral, because the amount of base neutralized by a

certain weight of one acid is also neutralized by an equivalent weight of another acid. It also follows from Richter's law that if one metal is precipitated by another metal from a neutral salt, the neutrality is maintained. Richter's generalization has been styled "the law of proportionality," "the law of equivalent ratios," etc., and it is obviously a special case of the law of reciprocal proportions of p. 42 recognized a few years later.

Neutralization versus hydrolysis.—The process of neutralization of a basic hydroxide by an acid is attended by the formation of a salt and water. We shall find later that some salts—*e.g.* zinc sulphate, sodium carbonate, potassium cyanide, etc.—are partially decomposed—*i.e.* hydrolyzed—by water into acid and base. The action of water on such a salt, or of an acid on such a base is thus an example of an opposing reaction; hydrolysis is opposed to neutralization:



In some cases, however, the amount of hydrolysis is inappreciable, and the process of neutralization is so complete that it can be employed for measuring the quantity of acid or base in a given solution. Here is a case where the influence of the solvent water is of prime importance.

Acidimetry and alkalimetry.—A standard solution containing a known amount of acid or base per litre is prepared, and just sufficient of this solution is added to neutralize a solution of a given base or acid. The volume of the standard solution required for the purpose is noted. It is possible to calculate the amount of "chemically pure" substances present in the given solution from the volume of the standard solution required for the neutralization. A standard solution containing one equivalent weight of the acid or base in grams per litre is called a normal solution, written "N-solution," and a solution containing one-tenth the concentration of a normal solution is called a decinormal solution, written " $\frac{1}{10}$ N-solution." The equivalent weight of a base is that quantity which just completely neutralizes one molecular weight of a monobasic acid; and the equivalent weight of an acid is that quantity which contains unit weight of replaceable hydrogen. Thus 36.47 grams of HCl per litre gives a normal solution of hydrochloric acid; and 49.04 grams of H_2SO_4 per litre gives a normal solution of sulphuric acid. Here the molecular weight of the acid is 98.08, and the acid is dibasic, for it contains two replaceable hydrogen atoms; and by definition:

$$\text{Equivalent of acid} = \frac{\text{Molecular weight of acid}}{\text{Basicity of acid}}$$

that is, the equivalent of sulphuric acid is $98.08 \div 2 = 49.04$. A normal solution of sodium hydroxide contains 40 grams of NaOH per litre, and a litre of a normal solution of any acid so far considered will just neutralize a litre of a normal solution of any base.

EXAMPLE.—Suppose that a 50 c.c. burette be charged with a normal solution of sodium hydroxide, and suppose that the amount of HCl in 500 c.c. of a dilute solution of hydrochloric acid be in question—acidimetry—pipette 50 c.c. of the acid into a beaker and add a few drops of litmus. The alkali solution is run from the burette into the acid in the beaker until the addition of but one more drop of alkali is needed to change the red litmus to blue. Suppose that 42 c.c. of the normal sodium hydroxide has been run from the burette. The argument runs: The neutralization $\text{NaOH} + \text{HCl} = \text{NaCl} + \text{H}_2\text{O}$ shows that 40 grams of sodium

hydroxide corresponds with 36.47 grams of HCl; and 1000 c.c. of NaOH has 40 grams of sodium hydroxide, which is equivalent to 36.47 grams of HCl. Consequently 42 c.c. of the standard sodium hydroxide solution is equivalent to 1.53 grams of HCl per 50 c.c. of the given acid or 15.3 grams of HCl are present in 500 c.c. of the given acid.

Similar remarks apply to the determination of alkalis—alkalimetry—by standard solutions of the acids. This process of analysis is called *volumetric analysis* in contradistinction to *gravimetric analysis*, which involves several weighings during each determination. In volumetric analysis, the substance taken for analysis is either weighed or measured, and the preparation of the stock of standard solution may involve one or two weighings; the stock of standard solution may last a great number of analyses. Experimental details are discussed in laboratory text-books.

§ 11. Bases.

We maintain that hydrogen is an essential constituent of all acids and alkalis.—J. P. COOKE.

A base—Greek *βᾶσις* (basis), a base—is a substance which reacts with an acid to produce a salt and water. For instance, zinc oxide reacts with sulphuric acid to produce zinc sulphate and water: $\text{ZnO} + \text{H}_2\text{SO}_4 = \text{H}_2\text{O} + \text{ZnSO}_4$. Sodium hydroxide reacts with sulphuric acid to produce sodium sulphate and water: $2\text{NaOH} + \text{H}_2\text{SO}_4 = 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4$. The bases include the oxides and the hydroxides of the metals, and certain groups of elements equivalent to a metal. For convenience, certain groups of elements like ammonia— NH_3 , hydroxylamine— NH_2OH , hydrogen phosphide— PH_3 , etc., are called bases, although they form salts by direct addition or combination without the separation of water. Thus ammonia and hydrogen chloride form ammonia chloride: $\text{NH}_3 + \text{HCl} = \text{NH}_4\text{Cl}$.

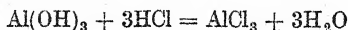
The definition of a base indicated above is highly unsatisfactory because it involves the definition of an acid, and we have just acknowledged that a satisfactory definition of an acid is not yet possible. Hence our definition of a base defines the unknown in terms of the unknown—*ignotum per ignotius*.

The term base was originally intended to express the ideal that the metal or metal oxide was the more important constituent, the foundation or base, so to speak, of a salt. This idea was dropped when it was recognized that the acidic constituent of a salt is just as important as the basic constituent. The idea persists in chemistry text-books when the salts are described under the basic element when usually they can be more concisely described with the acid component. However, I know of no book uniformly consistent in this respect.

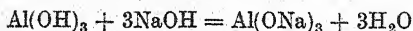
“Alkali” and “base” are not synonymous terms. Every alkali is a base, but every base is not an alkali. The alkaline oxides form very soluble hydroxides with marked basic properties. The oxides of the alkaline earths form sparingly soluble hydroxides with less marked basic properties. The other oxides, as a rule, do not react directly with water, and the hydroxides are made indirectly. An oxide cannot be classed as acidic or basic unless it can be shown to produce corresponding salts.

Peroxides.—We have seen how barium oxide, BaO —barium bivalent—when heated under certain conditions forms barium peroxide— BaO_2 . The peroxides contain a higher proportion of oxygen than the normal oxides. Barium oxide with sulphuric acid forms barium sulphate and water: $\text{BaO} + \text{H}_2\text{SO}_4 = \text{BaSO}_4 + \text{H}_2\text{O}$. It is therefore a base. Barium peroxide forms barium sulphate, water, and oxygen with sulphuric acid: $2\text{BaO}_2 + 2\text{H}_2\text{SO}_4 = 2\text{BaSO}_4 + 2\text{H}_2\text{O} + \text{O}_2$. $\text{Ba}(\text{SO}_4)_2$, not BaSO_4 , corresponds with barium peroxide. Hence, barium peroxide is not a basic oxide. In view of barium persulphate, $\text{Ba}(\text{SO}_4)_2$, barium peroxide becomes, by definition, a basic oxide.

Amphoteric oxides.—Lead dioxide or peroxide, PbO_2 —lead quadrivalent—is a basic oxide because it forms the corresponding salt— PbCl_4 . But PbO_2 also forms salts—plumbates—with bases, *e.g.* potassium plumbate, $\text{O} = \text{Pb} = (\text{OK})_2$. Hence, a substance may be both acidic and basic according to circumstances. Aluminium hydroxide—aluminium trivalent—is a base, because, when treated with an acid, it forms a salt—aluminium chloride, AlCl_3 —and water:

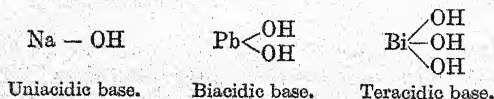


But aluminium hydroxide when treated with a base, say, sodium hydroxide, also forms a salt—sodium aluminate, $\text{Al}(\text{ONa})_3$ —and water:



Hence aluminium hydroxide acts towards an acid like a base, and towards a base like an acid. Such oxides can be called **intermediate oxides**, or **amphoteric oxides**—from the Greek *ἀμφότερος* (amphoterós), both. Zinc oxide is an intermediate oxide. Stannic oxide, $\text{O} = \text{Sn} = \text{O}$ —tin quadrivalent—forms stannic sulphate, $\text{SO}_4 = \text{Sn} = \text{SO}_4$, and also sodium stannate, $\text{O} = \text{Sn} = (\text{ONa})_2$; hence, stannic oxide is also an intermediate oxide.

Basic salts.—On comparing the graphic formulæ of the hydroxides of sodium (univalent), lead (bivalent), and bismuth (trivalent):



with the graphic formula for mono-, di-, and tri-basic acids we naturally inquire if the OH or hydroxyl group can be replaced by acid radicles one by one so as to furnish uni-, bi-, and ter-acidic bases. In the particular examples just selected, the salts:

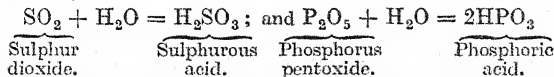


are known. The former is called basic lead nitrate, the latter normal lead nitrate, or simply lead nitrate. Similarly, $\text{Bi}(\text{OH})_2\text{NO}_3$, basic bismuth nitrate, is known. The basic salts are usually, not always, less soluble in water than the corresponding normal salts.

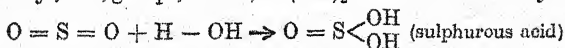
§ 12. Hydroxides and Anhydrides.

Acids and alkalis are compounds having the same general molecular structure and the differences between acids and alkalis, and we might add, the differences between individual acids and individual alkalis depend on the nature of the radicles.—J. P. COOKE.

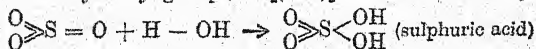
We have seen that sulphur dioxide and phosphorus pentoxide form acids with water:



The water in these compounds has completely lost its identity, and it is generally supposed to produce a new class of bodies called **hydroxides**. Every element, excepting fluorine and the argon family, appears to form one or more hydroxides, directly or indirectly. The oxides from which the acids are produced do not contain the elements of water. They are called **anhydrides**, or "acid anhydrides"—from the Greek *an*, without; *hydro* (hydor), water. Thus SO_2 is not only called sulphur dioxide, but also sulphurous anhydride; and P_2O_5 is not only phosphorus pentoxide, but phosphoric anhydride. An anhydride can be regarded as the residue left when the elements of water are removed from the oxyacids. Thus sulphuric acid, less water, gives sulphuric anhydride, SO_3 , also called sulphur trioxide; sulphurous acid, less water, gives sulphurous anhydride, SO_2 . It is generally supposed that sulphurous anhydride in combining with water forms a compound containing quadrivalent sulphur and two univalent hydroxyl, OH, groups, that is, $\text{SO}(\text{OH})_2$. The reaction is symbolized:

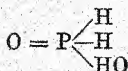


and sulphuric acid is considered to be a compound containing sexivalent sulphur and two hydroxyl groups, $\text{SO}_2(\text{OH})_2$. The reaction is symbolized:

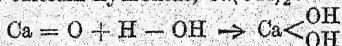


This subject will be resumed later.

Thus an anhydride with water forms an acid, and with a base it forms a salt: $\text{ZnO} + \text{SO}_3 = \text{ZnSO}_4$ (zinc sulphate).—Sulphurous acid can also be regarded as sulphurous hydroxide, $\text{SO}(\text{OH})_2$; and phosphoric acid—phosphorus quinquevalent—as phosphoric hydroxide, $\text{PO}(\text{OH})_3$. The basicity of an acid is generally supposed to correspond with the number of hydroxyl groups it contains. The hydrogen of the hydroxyl groups is supposed to be the "displaceable hydrogen" referred to in the definition of acids. Monobasic hypophosphorous acid, H_3PO_2 , is supposed to be $\text{H}_2\text{PO}(\text{OH})$; or



The basic oxides are sometimes called "basic anhydrides," and they too form hydroxides with water, *e.g.* calcium oxide, CaO —calcium bivalent—with water forms calcium hydroxide, $\text{Ca}(\text{OH})_2$:



From this point of view water can be regarded as hydrogen hydroxide, $H - OH$, analogous with $K - OH$, potassium hydroxide, and $Na - OH$, sodium hydroxide. Excluding certain carbon compounds, the hydroxides of the non-metallic elements are usually acids, and the hydroxides of the metals are usually bases. The term "hydroxide" is generally reserved for compounds of the basic oxides with water; and the term "anhydride" is usually reserved for the acid anhydrides.

Questions.

1. Oxygen is to be obtained from mercuric oxide, or manganese dioxide, or potassium chlorate. What is the cheapest method of making oxygen per litre, so far as raw materials are concerned, if 100 grams of mercuric oxide costs 9d.; 1 kilo of manganese dioxide, 10d.; and 1 kilo of potassium chlorate costs 1s. 3d.?

2. What is the meaning of the term "normal salt"? Why is the word "neutral" no longer used to designate these salts?

3. What volume in litres of oxygen measured at 730 mm. and 10° can be obtained from one kilogram of pyrolusite containing 80 per cent. of manganese peroxide?—*Edinburgh Univ.*

4. The atomic weight of carbon was formerly assumed to be 6, and of oxygen 8. State the grounds upon which these atomic weights have been doubled.—*Board of Educ.*

5. How do the acid-forming oxides (anhydrides) differ from the basic oxides? What is usually the effect of bringing together solutions of these two classes of oxides?—*G. H. Bailey.*

6. Give examples showing that the same oxide may at one time function as the acidic constituent of a salt and at another time as the basic constituent.—*G. H. Bailey.*

7. Explain what is meant by a "catalytic agent," taking as illustration the preparation of oxygen from potassium chlorate and manganese dioxide.—*Cape Univ.*

8. What do you understand by the term "a salt"? Give examples of normal, acid, and basic salts.—*London Univ.*

9. Define the following terms, giving examples in each case: acid, base, acid oxide, basic oxide, basicity, normal salt, acid salt, basic salt.—*Sydney Univ.*

10. There are reasons to believe that the equations: $2KClO_3 \rightarrow KCl + KClO_4 + O_2$, and $KClO_4 \rightarrow KCl + 2O_2$ do not actually represent the decomposition of potassium chlorate when heated. Describe the experiments you would make in order to test the validity of the equation.—*New Zealand Univ.*

11. The atomic weight of a metal M is 56. An oxide of this metal is found to consist of 53.8 per cent. of metal, and 46.2 per cent. of oxygen. Calculate the formula of the oxide.—*London Univ.*

12. A mixture of 5 vols. of hydrogen and 3 vols. of oxygen is exploded by an electric spark; will any gas remain? If so, how much, and how would you ascertain what it is? Suppose a mixture of 3 vols. of hydrogen and 5 vols. of oxygen had been used, what difference in the results would that have made?

13. Explain the following terms: oxide, hydroxide, and hydrate. Give examples of each. What do you understand by a peroxide and a suboxide?—*Science and Art Dept.*

14. Give the atomicity (valency) of each element in the following compounds: HCl , H_2O_2 , NH_3 , CO_2 , SO_2 , and SH_2 .—*Science and Art Dept.*

15. How much potassium chlorate must be heated to yield as much oxygen as would be obtainable from 500 grms. of mercuric oxide ($K = 39$, $Cl = 35.5$, $O = 16$, $Hg = 200$)?—*Board of Educ.* When two or more reactions follow in succession, or when two reactions are under comparison, it is not usually necessary to make a separate calculation for each reaction. Thus, since $KClO_3 \rightarrow KCl + 3O$, and $3HgO \rightarrow 3Hg + 3O$, it follows that $3HgO$ gives the same amount of oxygen as $KClO_3$. Ansr. 94.5 grms.

16. Why was oxygen so called? What objections can be urged against the name?—*London Univ.*

17. It is sometimes said that chemically a metal can be defined as an element capable of forming a base by union with oxygen. Discuss this definition.

CHAPTER IX

WATER

Molecular weight (vapour), $H_2O = 18.016$; (liquid) $(H_2O)_n = n18.016$. Melting point, 0° at atmospheric pressure; boiling point, 100° at atmospheric pressure; critical temperature, $+358^\circ$. Specific gravity of ice at 0° , 0.9167; and of water at 0° , 0.99987. One litre of steam at 100° and normal pressure weighs 0.590 gram.

§ 1. Water—Occurrence and Purification.

I am the daughter of the Earth and Water,
And the nursling of the Sky:
I pass through the pores of the ocean and shores;
I change, but I cannot die.—*The Cloud.*

Occurrence.—It has been estimated that three-fourths of the surface materials on the crust of the earth is water. Animals and plants contain a large proportion of combined water—*e.g.* fish contains the equivalent of about 80 per cent.; beef, 60–62 per cent.; the human body, 70 per cent.; aquatic plants between 95 and 99 per cent.; and ordinary land plants, 50–75 per cent. A great many rocks contain water—combined and absorbed. Clay, for example, contains up to 14 per cent. of combined water.

The circulation of water in nature—the water cycle.—Water is widely distributed in its three states of aggregation—steam or aqueous vapour, liquid water, and solid ice and snow. All the water on the earth passes through a remarkable cycle of changes. The heat of the sun leads to the evaporation of water from the seas, etc.; water vapour is only 0.62 times as heavy as an equal volume of air, and consequently it rises into the upper regions of the atmosphere, as well as diffuses into and mixes with atmospheric air. The temperature of the ascending vapour gradually decreases, and consequently a plane must be reached where the air is saturated with moisture. The vapour will then condense in the form of fine drops of water—mist or clouds. The fine drops coalesce into larger drops. Ultimately, the condensed water must descend again to the earth as rain, snow, or hail. The wind distributes the vapour. The heat given up during the condensation of the vapour is distributed or carried from the hotter regions—where evaporation is fastest—to the colder regions—where the vapour is condensed—thus helping to “stretch” the temperate regions nearer to the poles. The water which is sprayed, as rain, etc., on the surface of the globe, does a certain amount of physical and chemical work. On the chemical side, water helps in the decomposition and weathering of rocks; and on the physical side, it transports

matter in suspension from the higher to the lower levels. The soluble matters ultimately collect in the seas.

Thus the water cycle involves: (1) evaporation from the oceans, seas, lakes, etc.; (2) condensation in the upper regions of the atmosphere as a fine mist of distilled water where it collects as clouds; (3) further condensation is followed by rain; (4) the rain-water percolates through the soil and collects on an impervious bed of rock to be again forced to the surface, as spring water, by the pressure of the superincumbent layers; and (5) this is collected by the streams and rivers. The rivers return whence they came—to the sea, and commence anew the never-ending cycle. It must be added that a relatively small proportion of the water which finds its way into the ground falls out of the cycle since it is fixed by reaction with certain silicates and rocks forming hydrated silicates, hydrated aluminosilicates.

Rain-water.—"No water," said T. Bergman, "is ever found on the surface of the earth in a state of purity." Rain, in its journey through the air, dissolves oxygen, carbon dioxide, and nitrogen, as well as ammonia and ammonium nitrate. It also carries down dust—organic and inorganic. Rain water, particularly if collected near the sea in high winds, contains sodium chloride; and if collected near towns, sulphur compounds—sulphur dioxide and sulphuric acid—derived from the products of combustion of coal. When evaporated to dryness, 10,000 parts by weight of rain-water will give about 0.34 part of solid matter, most of this consisting of sodium chloride and organic matter. Rain-water contains in solution about 0.013 per cent. of dissolved nitrogen, 0.0064 per cent. oxygen, and 0.0013 per cent. carbon dioxide. The rain which falls at the end of a shower is more pure than that which falls at the beginning, because the atmosphere is washed, so to speak, during the earlier part of the shower.

Spring and mineral water.—Directly the rain-water strikes the ground, it begins to attack and dissolve various rocks, decaying organic tissue (humic compounds), etc., forming *surface and ground water*. It is estimated that between 25 and 40 per cent. of the rainfall, in temperate regions, soaks into the ground. In its journey underground—*underground water*—the percolating water loses most of its organic matter and dissolves more or less mineral matters—compounds of calcium, magnesium, and sodium—and also dissolves carbon dioxide, etc. The greater the depth to which the water sinks the greater the amount of solid matter it can dissolve. Water under great pressure is a powerful solvent. Sooner or later the water which has percolated underground will be forced to the surface as *spring water*. If the spring water holds an unusual amount of some particular constituent in solution which gives it a marked taste, or some specific property, the term *mineral water* is applied. Mineral waters do not necessarily contain a large excess of mineral matters in solution. The water from mineral springs is often named according to some special constituent dissolved in the water, or from the locality of the spring. "Fresh water" is a vague term applied to a natural water which does not contain much dissolved impurity.

Chalybeate waters contain ferrous carbonate—e.g. Tunbridge; Buxton; "Excelsior Spring," Saratoga, N.Y.; "Hot Springs" of Arkansas; Homberg; etc. *Sulphur waters* contain hydrogen sulphide and other sulphur compounds, alkaline sulphides, etc.—e.g. Baden; Harrogate; Bath; Aachen; "Red Sulphur

Spring," Sharon, N.Y.; etc. The water of the Steamboat Springs in Nevada has borates and deposits a sinter containing arsenic, antimony, mercury, lead and copper sulphides, as well as traces of gold and silver. *Saline waters* contain salts of various kinds, for instance, magnesium sulphate and chloride which give the water a bitter taste—e.g. Bath; Epsom; Seidlitz; Friedrichshall; Ofen; Cheltenham; etc. *Sodium sulphate and sodium carbonate*—e.g. Marienbad; Carlsbad; etc. *Carbon dioxide (acid reaction)*—e.g. Apollinaris (imitations of this and other mineral waters are made artificially; the natural water is bottled and exported). *Carbon dioxide with sodium carbonate (alkaline reaction)*—e.g. Vichy; Neuenahr; etc. *Carbon dioxide with sodium carbonate and sodium chloride*—e.g. Ems; Nieder-Selters; etc. *Sodium and other chlorides*—e.g. Homburg; Aachen; Baden-Baden; "Congress Spring," Saratoga, N.Y.; etc. *Iodine and bromine compounds*—e.g. "Congress Spring," and "Excelsior Spring," Saratoga, N.Y.; Woodhall Spa; etc. *Arsenic*—e.g. Roncigno; Levico; etc. *Lithia*—e.g. "Congress Spring," Saratoga, N.Y.; etc. *Boric acid*—q.v. *Silica*—q.v. *Hard waters* are described under "carbonates." The waters of some springs, particularly in volcanic districts, issue at an elevated temperature, hence the term *thermal waters*—e.g. "Hot Springs," New Zealand (about 60°); Teplitz (39°–49°); Vichy (32°); San Bernadine, California (40°–78°); etc.

River water.—Spring water collects in rivers and streams. It has been estimated that the rivers of the world discharge some 6524 cubic miles of water into the ocean per annum. Rivers contain not only the solid matter furnished by spring waters, but also organic matter derived from plants growing on the sides and bottom of the river, and also from the drainage of villages and towns through which the river passes. The river, in virtue of its greater volume and force, carries along a considerable amount of suspended solids. River water also contains in solution matter from the country which it drains. Thus the water of the Dee (Scotland), draining slate and sandstone rocks, contains about 0.0056 per cent. of solid matter, about one-fourth of this being calcium salts; the Thames, draining chalk rocks, contains about 0.03 per cent. of solid matter, two-thirds of which is calcium salts.

Sea-water.—Just as spring water flows into the rivers, the rivers flow into the sea carrying their dissolved salts, and suspended matters which have not been deposited in transit. The vapour which rises from the sea by evaporation is almost pure water; hence, unless the dissolved matters are continuously removed, sea-water must be gradually getting more and more "salty." Sea-water contains a relatively large proportion of soluble salts¹—about 3.5 per cent. of solids in solution. Where the evaporation is greatest we naturally expect to find the greatest proportion of salts in solution. Thus the Mediterranean contains about 3.4 per cent. of solids in solution; whereas the Baltic, with its numerous tributaries, and less evaporation, contains between 0.3 and 0.8 per cent. of solids in solution. Salts accumulate in land-locked seas and lakes much faster than in the sea. In illustration, the Dead Sea contains 22.8 per cent. of solids in solution; the Great Salt Lake (Utah), 23 per cent.; and the Elton Lake (Russia), 27 per cent. These masses of water behave as if they were exposed in a large evaporating basin, for the salts accumulate in the water and are deposited in crystalline masses on the shores of the lakes as the water evaporates.

¹ For instance, an average type contains about 98.5 per cent. of water; 2.7 per cent. of sodium salts; 0.07 per cent. of potassium salts; 0.14 per cent. of calcium salts; 0.59 per cent. of magnesium salts, as well as dissolved gases; 0.012 per cent. of nitrogen; 0.006 per cent. of oxygen; and 0.017 per cent. of carbon dioxide.

Potable and drinking water.—Potable water, that is, water which is fit for human consumption, is obtained principally from rivers and lakes, and also from wells—artesian and otherwise. The inorganic or mineral matters usually found in solution in natural water are not directly injurious to health. The purification of water for towns and cities is a very important practical problem for the chemist. The best solution can only follow after a careful study of the local conditions. Water should be free from pathogenic (disease-producing) bacteria, and from suspended impurities. This is generally done by filtration through large filter beds made from layers of sand and gravel, extending in some cases over an acre of ground. A Pasteur-Chamberlain's bougie (candle), made of unglazed and porous earthenware, and shaped like a hollow candle, is arranged to be screwed on to the water tap, Fig. 52. The water is forced through the earthenware by the pressure of the main. Bacteria, organic matter, etc., collect on the inside of the bougie as a slimy layer which clogs the filter. The bougie, Fig. 52, must be frequently cleaned or replaced (1) to permit

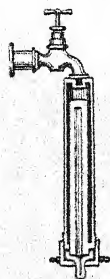


FIG. 52.—Filtration by Pressure.

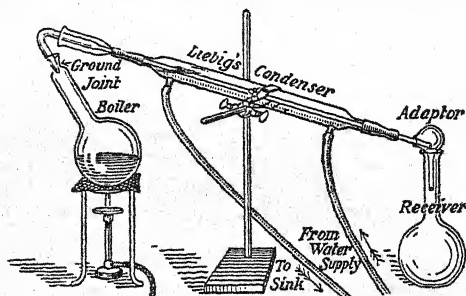


FIG. 53.—The Distillation of Water.

the free passage of water; and (2) to remove the layer of slimy organic matter which serves as a medium for the growth of bacteria. In some cases the bacteria are killed by the addition of minute traces of an oxidizer—ozone, sodium hypochlorite, copper sulphate, etc. The salt last named also prevents the growth of green algæ which are sometimes very troublesome.

To maintain the purity of the water supply up to the proper standard, it is necessary to make (1) a periodical critical examination of the source from which the water is obtained; (2) regular bacteriological examinations for pathogenic germs; and (3) chemical examinations for nitrogenous organic matter—albumenoids, etc.—upon which bacteria feed, and for the products of bacterial life—free ammonia, ammonium nitrate and nitrite. The presence of these substances in water throws it under suspicion.

The distillation of water.—Water is purified on a small scale by distillation. The water is boiled in a flask or boiler, and the steam is condensed back to the liquid condition by passage through a tube about which a continual stream of cold water flows. To economize space, the condensing tube is generally coiled as a spiral—called "the worm"—and kept in a tank through which cold water continually flows. Fig. 53 represents the form used by T. W. Richards in some atomic weight determinations.

It is a modification of that sometimes employed in the laboratory for distilling small quantities of liquid. The purpose of Richard's scheme is to prevent the steam coming in contact with rubber or cork stoppers, etc.—nothing but glass. In place of the “ground joint” a one-hole rubber or cork stopper and glass tube are often used; and in place of the “adapter,” the receiver is tilted so that the condenser tube dips into the neck of the flask. A small amount of volatile organic matter is usually carried over with the first rush of steam, and soluble matters derived from the glass may be found in the distillate. By adding a little potassium permanganate in alkaline solution to the water to be distilled, and rejecting the first portions carried over, fairly pure water can be obtained—particularly if the final product be redistilled in a platinum or gold vessel. Block-tin condensers are better than glass, since the water acts very much more vigorously on glass than it does on tin. Distilled water which has been kept some time in a glass bottle cannot be used satisfactorily in the analysis of silicates, because it contaminates the silicate undergoing analysis with some of the constituents to be determined.

§ 2. Some Properties of Water.

Water is absolutely indispensable to both animal and vegetable life; it is the cause of many of the most striking phenomena in nature; and it is employed for countless purposes by man.—P. F. FRANKLAND.

At ordinary temperatures, pure water is a tasteless and odourless liquid; it is colourless in moderately thin layers, but appears greenish-blue when viewed in thick layers. According to J. Aitken, the blue colour of large bodies of water—*e.g.* in china clay settling pits, in tanks in which water is being softened by the addition of milk of lime, etc.—is an optical effect due to the action of the fine particles suspended in the liquid on the light.

Liquids are but slightly compressible. If 1000 c.c. of water be subjected to a pressure of two atmospheres the volume will be reduced 0.05 c.c. According to P. G. Tait, this very small compressibility means that if sea-water were quite incompressible, the average level of the sea would be raised 116 feet higher than it is to-day, and 4 per cent. of the present land surface would be submerged.

Non-metallic liquids are bad conductors of heat; water is one of the best of liquids for conducting heat (mercury excepted), but even then, the thermal conductivity is small. Witness, a piece of weighted ice at the bottom of a test-tube of cold water. If the test-tube be held obliquely, and heated by a Bunsen's burner near the surface, the water at the surface will boil, but the ice at the bottom will remain unmelted.

Water boils at 100° under 760 mm. pressure. The greater the pressure, the higher the boiling point; and conversely, the less the pressure,¹ the lower the boiling point. These phenomena occur with liquids generally, and it is therefore necessary to state the pressure when giving the boiling point of a liquid. If no pressure is stated “760 mm.” is understood. Thus at Quito (9350 feet above sea-level), with the barometer at an average height, 525.4 mm. water boils at 90.1°; and on the top of Mount Everest (29,002 feet), barometer at 255.3 mm., water would boil at 72°. Steam

¹ Roughly about $\frac{1}{4}^{\circ}$ C. per mm. for a few degrees above and below 100°.

or water vapour is an invisible, colourless gas which condenses to a visible cloud of small particles when it comes in contact with the atmosphere. This is readily shown by boiling water in a flask; inside the flask, the vapour is invisible, and a cloud of minute water particles—condensed steam—appears where the steam comes in contact with the cold air.

Liquid water freezes at 0° into crystalline ice. Water vapour freezes into hoar frost and snow. The crystals of ice are extremely rare and difficult to measure. The crystals can often be seen when a piece of ice is examined with a lens while a beam of bright light is passed through it. Snow crystals are common. They appear in the form of an hexagonal (six-sided) nucleus or six-rayed star with the rays developed in bewildering complexity. The crystals are of inimitable delicacy and beauty. No two seem alike; but all are fashioned after one definite type—the six-rayed star—each one

Frail, but a work divine, made so fairly well,
So exquisite minute a miracle of design.—TENNYSON.

Few indeed would question what a writer has said; that if the chemist could wholly understand a drop of water he would know the origin and

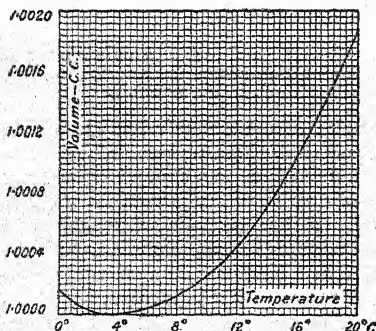


FIG. 54.—Relations between the Volume and Temperature of Water.

destiny of all things and hold the key to every happening. Allowing a due amount of poetic licence, the same remark could be applied to a drop of any other liquid or to a grain of any solid. Ice appears to be colourless or white when pure, but it is pale blue when seen in large masses.

By plotting the volume of a given mass of water at different temperatures, we get a curve similar to that illustrated in Fig. 54. This curve, at temperatures above 4° , shows that water, like most liquids, expands when heated and contracts when

cooled down to 4° ; but the curve below 4° is abnormal. It shows that water expands when cooled below, and contracts when heated up to 4° . If the specific gravity of water at 4° be taken as unity, it follows that water becomes specifically lighter when the temperature is raised or lowered beyond this point. The expansion of water when cooled from 4° to 0° is very small, but that minute quantity has a very important bearing in nature. When the water on the surface of, say, a lake is cooled, it contracts. The heavier cold water sinks, and the warm water rises. This circulation cools the temperature of the whole body of water down to 4° ; any further cooling results in the formation of specifically lighter water. Accordingly, this remains on the surface, and circulation ceases. Finally, as a result of this remarkable and abnormal property, when the temperature of the atmosphere falls to 0° , a surface film of ice is formed.¹ If the water did not expand in this way, as the

¹ "Ground ice" or "anchor ice" is formed at the bottom of rapidly moving streams when the water is thoroughly mixed and does not settle in layers.

temperature fell to 0° , the whole body of water would freeze from below upwards and produce profound climatic changes, since the larger amount of ice formed in winter would materially affect the temperature for the rest of the year. These remarks do not apply to sea-water which contracts as the temperature is lowered down to the freezing point. In the act of freezing water expands so that 100 c.c. of liquid water at 0° gives approximately 110 c.c. of ice at the same temperature. The specific gravity of ice at 0° varies with its mode of formation from 0.9159 to 0.9182; the specific gravity of water at 0° is 0.999867. Accordingly, ice floats on the surface of water. The expansion of water during freezing is an important factor. The expansion may burst the intercellular tissue of plants by freezing the cell-sap; the expansion may disrupt the fibres of flesh, so that frozen meat appears rather more "pulpy" than ordinary meat. If water freezes in pipes, the expansion of water in the act of freezing may burst the pipe, and water will "leak" when the ice "thaws"; water freezing in the surface crevices of rocks splits and widens the fissures so that the surface crust of the rock appears to disintegrate during a "thaw." The *débris* collects as "talus" at the foot of the rocks, ready to be transported by water to lower levels. Hence this simple force plays an important part in the weathering and decay of rocks, building stones, etc., in countries exposed to alternate frost and thaw; and J. Tyndall adds: "The records of geology are mainly the history of the work of water."

§ 3. The Molecular Structure of Water.

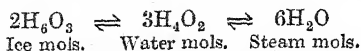
Not only in the matter of solutions, but in other more strictly physical relations, it is a misfortune that the rôle of the typical liquid was assigned to water.—G. F. STRADLING (1901).

Water is water—not a single substance in the proper acceptation of the term, but a "mush" or mixture; an entirely peculiar material not to be represented by any one formula, nor spoken of by a single name.—H. E. ARMSTRONG.

Steam or water vapour.—The student so frequently represents water by the formula H_2O , that he soon believes that this symbol correctly represents the molecule of liquid water. As a matter of fact, the molecule of water is probably much more complex. The vapour density of steam is rather too great for the molecular formula H_2O , and much too small for H_4O_2 . It is therefore assumed that steam contains a mixture of H_2O with a few H_4O_2 molecules, and that the equilibrium condition for water vapour, $H_4O_2 \rightleftharpoons 2H_2O$, corresponds with 91 per cent. of H_2O molecules in the vicinity of 100° .

Liquid water.—In order to explain the very curious physical properties of liquid water, e.g. Fig. 54, it was assumed by W. Röntgen (1891) that water is a mixture of two kinds of molecules which he called "ice molecules" and "water molecules." The "ice molecules" were supposed to form a mass more complex though less dense than "water molecules," so that during the fusion of ice and the subsequent warming of the ice-cold water, the volume contracts as indicated in Fig. 54. Later investigations have led to the assumption that liquid water contains three kinds of molecules, which are formed by the association of, say, n simple molecules of H_2O to form more complex molecules, $(H_2O)_n$. There may be some difference of opinion as to the numerical values of n , but there is little room

for doubt about the existence of complex molecules. The simplest assumptions are that $n = 1, 2$, and 3 , so that the three kinds of molecules are represented: H_2O , which has been called *hydrol*, or steam molecules; H_4O_2 , called *dihydrol*, or water molecules; and H_6O_3 , called *trihydrol*, or ice molecules.¹ Assuming these molecular weights are correct, it is assumed that the peculiarities in the physical properties of water, and indeed also the chemical properties, can be explained in terms of the balanced reactions



An increase of temperature or pressure favours the passage from left to right; and a decrease of temperature or pressure, the passage from right to left.

Ice or solid water.—G. Tammann and P. W. Bridgman, have studied the effects of great pressures on the properties of ice, and deduced the

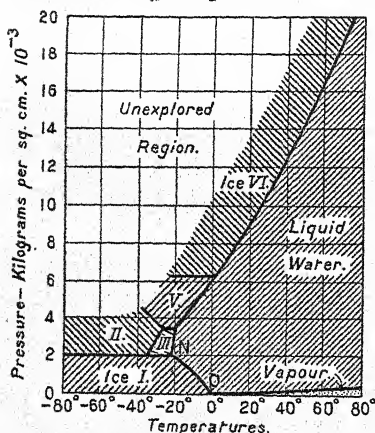
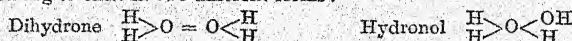


FIG. 55.—The Equilibrium Conditions of the different varieties of Ice.

existence of five distinct varieties: Ordinary ice, which crystallizes in the hexagonal system, and is called Ice I, or light ice, because it is lighter than water at ordinary temperatures. Ice III, or dense ice, is formed by subjecting ice to pressures exceeding 2000 atmospheres, and it passes at once into ordinary "Ice I" if the pressure be reduced. The transformation of "Ice III" into "Ice I" can be delayed by reducing the temperature, so that samples can be examined in vessels at the temperature of liquid air, about -180° . If dense ice be made by freezing water while under a pressure of about 3000 atmospheres, a variety of dense ice—called Ice II—is obtained which is not the same as "Ice III," made by compressing ice itself to 3000 atmospheres. "Ice III" is stable below 33.4° at 2220 atm. pressure, and below -26° at 3116 atm. pressure. A fifth variety, Ice V, has been prepared at -17° under a pressure of 3420 atm. This passes into a sixth variety, Ice VI, at 6170 atm. pressure. Ice VI exists only above 0°C ., and the effect of pressure on its melting point has been traced up to the comparatively high temperature 76.35° at 20,000 atm. pressure. The existence of another variety, Ice IV, announced by G. Tammann, was

¹ In an attempt to account for the properties of dilute solutions, H. E. Armstrong (1908) further assumed that some of these molecules have the same composition, but a different structure; thus the dihydrol, H_4O_2 , molecule is assumed by Armstrong to exist in two different forms:



But this is not the place to discuss the evidence adduced in favour of a belief in the real existence of these imaginary substances.

not confirmed by P. W. Bridgman. The results of the latter are shown graphically in Fig. 55; and the specific gravities of the five forms are 0.92 for Ice I; 1.03 for Ice II; 1.04 for Ice III; 1.06 for Ice V; and 1.09 for Ice VI.

If these views about the constitution of steam, liquid water, and ice be valid, it is hardly correct to say, without some reservations, that the passage of ice to liquid water and to steam, and the converse changes, are purely physical changes. Confer §§ 2 and 3, Chapter II.

§ 4. The Vapour-pressure of Water. Boiling.

So long as a body retains its normal state of aggregation and properties, we can observe an increase of temperature corresponding with an increase of molecular energy, but as soon as the destruction of form begins to take place, the increase of temperature no longer becomes sensible, and the energy is diverted to breaking up the structure of the body, and keeping its molecules apart. When this has been accomplished, and not till then, additional energy again produces additional motion, and the substance gets hotter and hotter.—W. ANDERSON.

Kinetic theory of evaporation.—The molecules of a liquid are probably much closer together than is the case with gases, and they are accordingly subjected to the action

of comparatively powerful inter-molecular forces. Diffusion also shows that the molecules of a liquid are in motion, but, in consequence of the great overcrowding, the number of collisions must be comparatively great. The molecules in the body of the

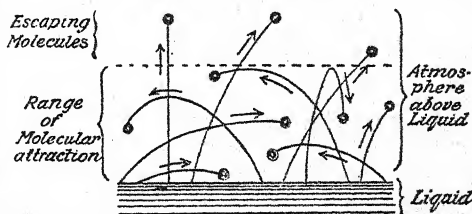


FIG. 56.—Diagrammatic Representation of the Molecules just above the Surface of an Evaporating Liquid.

liquid are attracted by the other molecules equally in all directions, but at the surface, the molecule can only be attracted inwards. What will happen to a molecule, which, in the course of its wanderings, reaches the surface? If its velocity be great enough, the molecule will rush upwards beyond the range of attraction of the other molecules in the liquid, and thus pass into the space above. On the other hand, if the velocity of the escaping molecule be not great enough to carry the molecule so far, the upward velocity of the molecule will become less and less, and finally the molecule will fall back and plunge into the liquid again. The case is somewhat analogous with the behaviour of a stone thrown into the air. If the stone were projected upwards with a sufficient velocity, say 50,000 feet per second, it would leave the earth never to return. Hence if the boundary surface of a liquid could be magnified sufficiently, and if the kinetic theory be correct, a cross-section above the surface of the liquid would present some such appearance as indicated in Fig. 56 (after A. D. Risteen). The trajectories of the molecules are also shown in the diagram. Many of the molecules which leave the surface of the water fall back again; and those which leave and do not return reduce the volume of the liquid, and finally lead to the complete evaporation of the liquid.

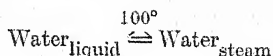
Raising the temperature of the liquid accelerates the motions of the molecules and so hastens the process of evaporation. A draught of air across the surface also favours the passage of the molecules away from the atmosphere above the evaporating liquid and reduces the chance of return.

Cooling effect during evaporation.—A liquid becomes cooler during evaporation. In 1755 W. Cullen placed water under a bell-jar from which the air was rapidly withdrawn, the evaporation was so rapid that the water was cooled until it froze; similarly by placing some liquid sulphur dioxide and water in a red-hot crucible, the water has been frozen therein, being cooled by the rapid evaporation of the sulphur dioxide. The kinetic theory shows how this can occur. During evaporation, the fleetest molecules can alone escape from the liquid, the more sluggish molecules cannot get beyond the range of attraction of the molecules remaining in the liquid. The fleetest molecules have the greatest kinetic energy, and we have seen, p. 140, that the temperature of a mass of molecules is proportional to the average kinetic energy of the molecules. If, therefore, the fastest molecules escape, the more sluggish molecules will remain behind, and the average velocity of the molecules of the liquid must be reduced. Hence a liquid which is evaporating is cooling rapidly. To illustrate the cooling effect of evaporation, a little ether is placed in a small beaker with a few drops of water on the underside, the water will freeze if the ether be evaporated quickly by blowing a jet of air across the surface. Advantage is taken of this fact to solidify carbon dioxide by the rapid evaporation of liquid carbon dioxide, to solidify hydrogen by the rapid evaporation of liquid hydrogen, etc. The principle is also utilized in cold storage, etc.

Heat of evaporation and fusion.—A relatively large amount of energy is needed to transform a gram of water into steam. The thermal energy required has to accelerate the motion of the individual molecules (specific heat) and at the same time to impart to the molecules sufficient momentum to tear them apart against molecular attraction (latent heat of vaporization). The amount of heat required to turn one gram of water at 100° into steam at 100° is 537 cal.¹ This is the latent heat of vaporization of water. The number means that steam at 100° has the equivalent of 537 cal. of energy—internal or potential—more than liquid water at 100° . This energy is degraded as heat when steam at 100° is cooled to liquid water at 100° . Similar remarks might be applied to the melting of ice to liquid water; and the freezing of liquid water into ice. In this case, one gram of ice at 0° in melting to liquid water at 0° requires about 80 cal.—this is the so-called latent heat of fusion. Similar relations hold for other substances. If secondary changes—e.g. decomposition during fusion—do not occur, all substances exhibit characteristic latent heats of fusion and vaporization. In virtue of these facts, it follows that weight for weight a liquid contains a greater amount of energy than a solid, and a gas contains a greater amount than a liquid. In order to change a solid to a liquid, or a liquid into a gas, energy must be added to the substance, and for the converse changes, gas to liquid, or liquid to solid, energy must be withdrawn from a substance. The energy needed for the evaporation of natural waters is mainly derived from the "heat paid out by the sun."

¹ A calorie, by the way, is a unit of heat represented by the amount of heat required to raise the temperature of one gram of water one degree.

Vapour pressure.—Suppose that a liquid is evaporating in a closed vacuous space. The fleetest molecules cannot escape into boundless space, and consequently they accumulate as a gas or vapour in the space above the liquid.¹ The concentration of the vapour in the space above the liquid will go on increasing. The molecules of the vapour behave like the molecules of an ordinary gas, and consequently a certain percentage will plunge back into the liquid. The number of molecules which return to the liquid from the space above per second of course increases as the concentration of the vapour increases, although the rate at which the molecules leave the liquid probably decreases as the concentration of the vapour increases. When the number of molecules which return to the liquid in a given time is equal to the number of molecules which leave the liquid in the same time, the vapour is said to be saturated, and the system in equilibrium. With the notation previously used :



The equilibrium, it will be observed, is not a static condition, that is, a state of rest. Both processes are active (kinetic), not passive (static). There is a shower of molecules streaming into the liquid, and an efflux of molecules away from the liquid. The effect of one is neutralized by the other; neither can produce any visible result. Anything which disturbs this equality—e.g. a desiccating agent or a condenser in the space above (as in distillation, p. 180), etc.—will alter the conditions. Experiment shows that at a given temperature the vapour pressure of a liquid in contact with its own liquid is a constant quantity, and independent of the absolute amount of vapour and of liquid present in the system. It is easy to see this.

If the surface of the liquid be doubled, it is true that twice as many molecules will leave the surface in a given time, but twice as many molecules will return.

The vapour pressure of water at 0° is just equal to 4.60 mm. of mercury. This means that if a little water be introduced into the Torricellian vacuum of a barometer, at 0°, the mercury will be depressed from 760 mm. to 755.4 mm. The higher the temperature, the greater the vapour pressure, provided all the water is not vaporized; but for any assigned temperature the vapour pressure of a given liquid always has one fixed and definite value. The effect of temperature on the vapour pressure of water is indicated by the curve, Fig. 57, which

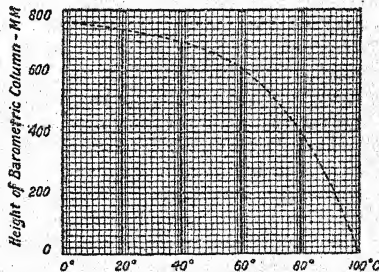


Fig. 57.—Vapour Pressure of Water.

¹ The distinction between "gas" and "vapour" is somewhat vague. If the "elastic fluid" be very far from its temperature of liquefaction, it is generally called a "gas"; and "vapour" if it is near its temperature of liquefaction. E.g. oxygen, nitrogen, etc., at ordinary temperatures are gases; whereas water or alcohol on evaporation would furnish vapours. Otherwise expressed, a gas is an elastic fluid at a temperature above its critical temperature, and a vapour is an elastic fluid below its critical temperature, but not in a liquid state.

shows incidentally the effect of introducing a thin layer of liquid water into the Torricellian vacuum; the diagram shows the height of the barometer at different temperatures. The vertical distances between the two dotted curves represent the vapour pressures of water at the corresponding temperatures.

Boiling or ebullition.—As we have just seen, raising the temperature of an evaporating liquid increases the average speed of the molecules, and favours rapid evaporation. When the temperature is high enough, the exposed surface of the liquid is not sufficient to allow the swift-moving molecules to escape fast enough, bubbles of vapour are accordingly formed *within* the liquid. Each bubble as it forms rises to the surface—increasing in size as it rises—and finally escapes into the atmosphere. The process of vaporization by bubble formation is called boiling; and the temperature at which boiling commences, the boiling point of the liquid. When the vapour pressure of the liquid is the same as the external pressure to which the liquid is subjected, the temperature does not usually rise any higher. Increasing the supply of heat increases the rate at which the bubbles are formed. Hence it is sometimes convenient to define: The boiling point of a liquid is the temperature at which the vapour pressure of the liquid is equal to the external pressure exerted at any point on the liquid surface. This external pressure may be exerted by the atmospheric air, by vapour and air, by other gases, etc. Hence a table of the vapour pressures of a liquid at different temperatures also shows the boiling points of that liquid under different pressures. Thus water at a pressure of 4.6 mm. of mercury boils at 0°. Hence liquids which decompose at their boiling point under ordinary atmospheric pressure can frequently be distilled without decomposition at the lower boiling temperature obtained by reducing the pressure. Hydrogen peroxide can be cited in illustration. Hence the so-called process of distillation under reduced pressure, or, as it is sometimes less accurately styled, distillation *in vacuo*. The process is illustrated later, Fig. 86.

It has been estimated that there are not less than 1,000,000,000,000,000,000 molecules per c.c. of air at ordinary temperature and pressure; when the pressure of the air is reduced to 0.1 mm., there will still be about 100,000,000,000,000 molecules per c.c.; at 0.01 mm. pressure, 10,000,000,000,000 molecules per cc. Hence in the best of the so-called vacua obtained by the air pump there must be left an immense number of molecules.

Measuring the volume of moist gases.—We have discussed evaporation in an evacuated space. In 1802 John Dalton showed that *the mass of vapour required to saturate a given space at a given temperature, and consequently also the vapour pressure of a given liquid, is the same whether the vapour be by itself, or associated with other gases upon which it has no chemical action.* In other words, the total pressure of a mixture of gas and vapour is the sum of the partial pressures of each constituent. Hence, added J. Dalton, "if the aerial atmosphere was at once annihilated, leaving only its aqueous vapour, this would be but little augmented because it already exists in air nearly at a maximum of that which the temperature can produce and support. The suppression of the obstacle will only accelerate the evaporation without sensibly augmenting the absolute quantity." When a gas is confined over water, the observed volume of the gas is determined by the temperature and pressure of the barometer. By Dalton's law

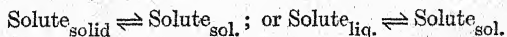
the total pressure of the gas is the joint effect of two partial pressures: (1) the partial pressure of the water vapour at the observed temperature; and (2) the partial pressure of the gas under observation. Hence the barometer reading does not represent the pressure of the gas alone, but rather the pressure of the gas plus the pressure of the water vapour. To find the latter, note the temperature, and a "Table of Vapour Pressures of Water in Millimetres of Mercury," in most "Table Books," will furnish the desired datum. This must be subtracted from the observed pressure (barometer) in order to find the pressure of the gas at the temperature in question.

EXAMPLE.—What is the volume of $4\frac{1}{2}$ litres of a gas at 0° , 760 mm., when it is measured in contact with water at 15° , and the barometer reads 767.7 mm.? From the "Table of Vapour Pressures," the vapour pressure of water at 15° is 12.7 mm. Hence the gas is really under a partial pressure of $767 - 12.7 = 755$ mm. of mercury. The problem is now to be solved like the example on p. 105. The answer is 4.24 litres.

§ 5. The Solubility of Solids.

In the strictly scientific sense of the word insolubility does not exist, and even those substances characterized by the most obstinate resistance to the solvent action of water may properly be designated merely as extraordinarily difficult of solution, not as insoluble.—O. N. WRR.

Water is one of the most active of solvents, and, in consequence, it has been styled "the universal solvent," but not in the same sense as the visionaries' *alcahest* (universal solvent) so often mentioned in mediæval alchemy. They seem to have overlooked that no vessel could be made to hold a universal solvent. For convenience, the dissolved substance is often called the solute, and the liquid in which the solute is dissolved the solvent. If potassium chloride be added to water kept at a constant temperature, the salt is gradually dissolved, and the process of solution continues until a definite amount has dissolved. The amount of solid remaining in excess of this will remain an indefinite time without farther change, provided the temperature remains constant, and no solvent is lost by evaporation. The solid and solution are then in equilibrium. As in the analogous case of the vapour pressure of a liquid, the equilibrium between a saturated solution and a solid is dynamic, not static. Accordingly, with the preceding notation the equilibria respectively with solid and liquid solutes are represented:



The solution is said to be saturated with the salt at the temperature of experiment. The weight of salt dissolved by 100 c.c. of the solvent so as to make a saturated solution at any assigned temperature is called the solubility of the salt.¹ Thus, 100 c.c. of water at 20° will dissolve 35 grams of potassium chloride, and accordingly, 35 is the solubility of potassium chloride in water at 20° . The concentration of a solution is determined by the relative amount of solute in solution. If but a small

¹ Other modes of representing solubility are more convenient in special cases—e.g. in Fig. 58 the percentage amount of salt in a given weight of the solution is employed; the volume of gas at n.p.t. per 100 volumes of solvent is used on p. 826; etc.

proportion is present, the solution is said to be *weak* or *dilute*; if a relatively large amount of solute is dissolved, the solution is said to be *strong* or *concentrated*. J. H. van't Hoff defined a solution as a homogeneous mixture of two or more substances whose composition can vary within certain limits; C. L. Berthollet used a similar definition. The limiting amount of salt required to saturate a solvent was called by N. le Febure (1660) *un poids de nature* which cannot be transgressed. For a discussion whether solutions are chemical compounds or not, see page 41.

The so-called insoluble substances obtained as precipitates in analysis are in reality substances with a very low solubility. It is all a question of measurement. As the methods of observation become more and more precise, so does the list of insoluble substances grow less and less. The general use of the term "insoluble" must in consequence give way to "sparingly soluble." In illustration, the three precipitates obtained in the first group of the regular scheme for qualitative analysis are not really insoluble in water because their solubilities, per 100 c.c. of water, at 20°, are represented by the following numbers: silver chloride, 0.00016 gram; mercurous chloride, 0.00031 gram; and lead chloride, 1.18 grams. In order to maintain Witt's proposition at the head of this section, it would be necessary to make so many logical leaps in the dark, that the proposition is a good illustration of a useless proposition not at all uncommon in chemistry—*vide* p. 124.

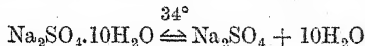
The influence of temperature.—The solubility of most solids increases with the temperature; the higher the temperature, the greater the solubility. The relation between the solubility of solids and temperature is usually plotted in the form of curves called *solubility curves*. The solubility curve presents a graphic picture which enables the relation between solubility and temperature to be seen at a glance. In illustration, Fig. 197 shows the solubility curves of potassium and sodium nitrates; Fig. 65 shows solubility curves of potassium chloride, chlorate, and perchlorate. The upward left-to-right slope of the curve shows that the solubility of these salts *increases* with a rise of temperature. Sodium chloride is but slightly more soluble in hot than in cold water, and a few substances give solubility curves with a right-to-left upward slope, showing that the solubility *decreases* with rising temperatures. For instance, see the curve *EO*, Fig. 58, for the solubility of anhydrous sodium sulphate; and also:

SOLUBILITY OF	0°	50°	100°
Calcium chromate— CaCrO_4	4.50	1.12	0.42
Calcium hydroxide— $\text{Ca}(\text{OH})_2$	0.14	0.10	0.08
Cerium sulphate— $\text{Ce}_2(\text{SO}_4)_3$	19.09	4.78	0.78
Sodium chloride— NaCl	35.63	36.67	39.12

Some solubility curves exhibit irregularities at certain temperatures. The solubility curve may change its direction, as calcium sulphate does at 35°, and barium butyrate at 45°. The solubility curve of sodium sulphate is a very trite illustration, but none the less instructive on that account. It is shown in Fig. 58¹. The solubility of the salt, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$, *increases*

¹ The ordinates and abscissæ in this diagram are usually reversed. It really makes no difference, but for the sake of ease I prefer the system used in the text, since we can then see the relation between solubility curves and temperature; concentration diagrams usually represented as in Fig. 59.

rapidly with rise of temperature, as shown by the slope of the curve *EO*, Fig. 58. There is an abrupt change in the direction of the solubility curve at 34°, Fig. 58. Above that temperature the solubility *decreases* with rise of temperature. The interpretation is that the solubility curve is really compounded of two solubility curves. The curve of increasing solubility with rise of temperature below 34° represents the solubility of curve of the decahydrate, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$; and the curve of decreasing solubility with rise of temperature represents the solubility curve of the anhydrous salt, Na_2SO_4 , in rhombic crystals. The decahydrate, at 34°, is transformed into the anhydrous salt. The decahydrate is not stable above 33°; the anhydrous salt is not stable below 34°. This temperature is called the transition temperature or transition point, and the change is symbolized



The solubility curves, it will be observed, represent the condition of equilibrium between the solvent and salt.

It makes no difference whether we start with the anhydrous sulphate or the decahydrate. When in equilibrium, the solution in contact with the solid will contain the amounts of sodium sulphate— Na_2SO_4 —indicated by the solubility curves, Fig. 58. The saturated solutions, when in equilibrium, have the same concentration and are identical in every way. We cannot continue the observation of the solubility of the decahydrate beyond 32.383° because it *immediately* splits up either into the anhydrous form, or into some less hydrated form—e.g. $\text{Na}_2\text{SO}_4 \cdot 7\text{H}_2\text{O}$; but since the transformation of the anhydrous salt into the hydrate takes an appreciable time, it is possible to measure the solubility of the anhydrous salt below 32.4. This is indicated by the dotted line in the diagram.

The solubilities of the two sodium sulphates—anhydrous and decahydrate—are quite different. At 234° the rhombic sodium sulphate passes into the monoclinic form with its own solubility curve. If the solid decahydrate were in contact with a saturated solution at 20°, and some of the anhydrous sulphate were added to the solution, some of the anhydrous sulphate would dissolve and be deposited later as the decahydrate. The final result would be a transformation, through the medium of the solution, of the anhydrous salt into the decahydrate. Although 100 c.c. of water at 0° can dissolve only about 5.0 grams of the decahydrate, the same quantity of water can dissolve much more of the anhydrous sulphate. It is therefore necessary to specify which sodium sulphate is in question when reference is made to "a saturated solution of sodium sulphate," unless "at ordinary temperature" is understood as is usually the case. The fact that salts

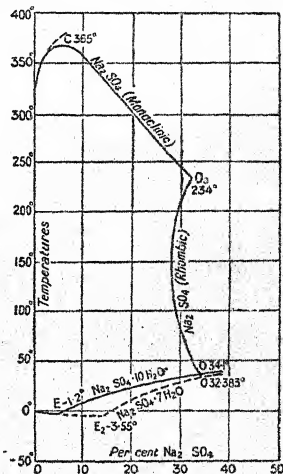


FIG. 58.—Solubility Curve of Sodium Sulphate.

which crystallize in two or more forms with different amounts of combined water have different solubilities agrees with the assumption that the solution itself contains molecules of the hydrate—*e.g.* $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$; and $\text{Na}_2\text{SO}_4 \cdot 7\text{H}_2\text{O}$.

§ 6. The Freezing of Solutions.

It has long been known, even as far back as Aristotle's day, that drinkable water could be obtained from frozen sea-water; and that if an aqueous solution of salt be gradually cooled, comparatively pure ice first separates from the solution. The trace of salt generally found in ice which has separated from sea-water is mechanically entangled in the ice. Magnesium cyanoplatinate, for instance forms a colourless solution from which colourless ice separates; if any solid magnesium cyanoplatinate separated with the ice, the colour would be intensely red.

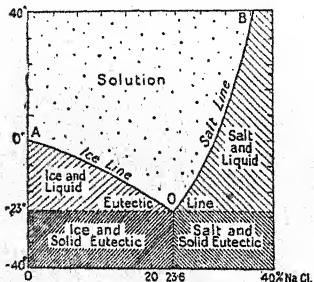


FIG. 59.—Solubility and Freezing Curves of Sodium Chloride-Water Solutions.

The curve *OB*, Fig. 59, represents the solubility of sodium chloride at temperatures ranging from -23° to $+40^\circ$; the observation cannot be continued below -23° because the whole mass freezes; the upward curve would probably stop only at the melting point of sodium chloride, 801° , if it were not for the volatilization of the water. Hence, to determine the solubility, the pressure would have to be very great at this high temperature to prevent the water volatilizing. Note the critical temperature of water.

The freezing temperature of a solution is generally lower than that of the pure solvent. More than a century ago, C. Blagden (1788) cited a number of observations which led him to the belief that the lowering of the freezing point is proportional to the amount of substance in solution. In Blagden's own words: The effect of a salt is to depress the freezing point in the simple ratio of its proportion to water. This generalization is sometimes called Blagden's law. The freezing point of an aqueous solution of sodium chloride, that is, the temperature at which ice begins to separate, is gradually reduced by the continued addition of small quantities of sodium chloride, and reaches its lowest value, -23° , when the solution has nearly 23.6 per cent. of sodium chloride; further additions of the salt raise the temperature at which solid separates. *Solid sodium chloride*, not ice, then separates from the solution. F. Guthrie's measurements (1875), of the relation between the freezing point and the concentration of aqueous solutions of sodium chloride are shown graphically by the curve *AO*, Fig. 59.

Imagine a 5 per cent. solution of salt subjected to a gradually diminishing temperature. Start at 0° . When the temperature reaches, say, -3.4° ice separates from the solution. The mother liquid remaining has therefore more than 5 per cent. of salt in solution; as the temperature falls, more ice separates. The further concentration of the mother liquid and

the separation of ice continue until the mother liquid has about 23.6 per cent. of salt, when the whole remaining liquid freezes *en bloc* at -23° . Quite an analogous sequence of changes occurs if solutions containing more than 23.6 per cent. of salt be gradually cooled. This time, however, instead of pure ice, pure salt separates until the residual liquid has 23.6 per cent. of salt. The whole solidifies *en masse* at -23° . If the cooling solution has just 23.6 per cent. of salt, neither ice nor salt separates, until the temperature has fallen to -23° , when the whole freezes to a solid mass. No other mixture of water and salt freezes at a lower temperature than this. Hence a solution containing 23.6 per cent. of salt is called a eutectic mixture;¹ and -23° the eutectic temperature. F. Guthrie used to think that this mixture—water with 23.6 per cent. of salt—corresponded with the formation of a definite compound of sodium chloride and water— $\text{NaCl} \cdot 10\text{H}_2\text{O}$ —stable only at low temperatures. Hence his designation cryohydrate for the alleged compound. Ponsot called the substance a cryosol. The term “eutectic mixture” is preferred in place of cryohydrate or cryosol.

Cryohydrates.—We now know that Guthrie’s cryohydrates are nothing but mechanical mixtures of ice and salt. The one is entangled with the other. It is easy to understand how eutectic mixtures were mistaken for true chemical compounds. No matter what the original composition of the salt solution, the last fraction to solidify always has the same composition; and a constant melting point. Both these qualities are often stated to be characteristics of true chemical compounds. That the cryohydrates of sodium chloride and numerous other salts are not chemical compounds is based on the following evidence: (1) The heterogeneous structure is frequently apparent under the microscope. The crystals of ice can often be seen lying in a matrix of salt, especially if a coloured salt like potassium permanganate, copper sulphate, or potassium dichromate be employed. See Fig. 195 for an illustration with solutions of carbon in iron. (2) Unlike true crystalline compounds, the cryohydrates are generally opaque and ill-defined. (3) Alcohol may dissolve the solvent, leaving behind a network of salt. (4) There are no special signs of chemical change during the formation of the cryohydrate. (5) The ratio of salt to solvent is not always in molecular proportions. The agreement in some cases is merely a coincidence. (6) The composition of a cryohydrate is different when the solidification takes place under different pressures.

The phenomenon—eutexia—which is illustrated above, has attained great importance in the study of metallic alloys, minerals, and mixed salts. Other examples will appear later. If a recording thermometer be placed in a cooling solution, say, sodium chloride, a series of curves like those indicated in Fig. 60 will be obtained. A pure liquid shows one break

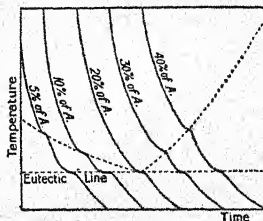


FIG. 60.—Cooling Curves of Binary Mixtures which do not form a Chemical Compound.

¹ From the Greek *eu* (eu), easily, and *τέκω* (teko), I melt. Hence “eutectic” means “that which is easily melted.” The word *εὐρηκτός* was used by Aristotle (*Problemata*, 1. 50) in the sense of easily soluble or digestible.

which corresponds with the latent heat of solidification liberated at the freezing point; a solution, corresponding with the eutectic mixture, also shows only one break at the eutectic temperature. Other mixtures show two breaks, one at which a solid separates and the other at the eutectic temperature. If the proportion of solvent is greater than corresponds with the eutectic mixture, the first break indicates the temperature at which the solvent begins to separate; and if the proportion of solvent is less than corresponds with the eutectic temperature, the first break indicates the temperature at which the solute begins to separate. A comparison of Figs. 59 and 60 will make this clear.

§ 7. Gibbs' Phase Rule.

The phase rule is one of the most comprehensive generalizations known to man. It is of unlimited application, and offers an accurate and ready means of classifying all states of physical and chemical equilibria.—W. MAYERHOFER.

On plotting the vapour pressure of water in the presence of its own liquid, we get the curve *OQ*, Fig. 61, which gives the vapour pressure of

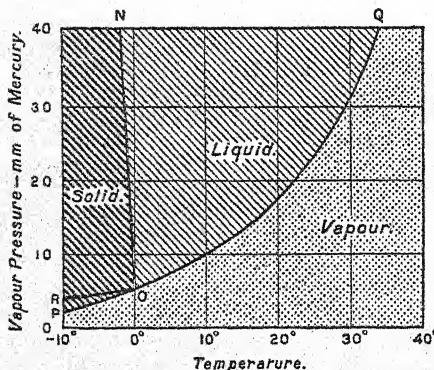


FIG. 61.—Vapour Pressure Curves of Water.

water corresponding with any given temperature when the liquid and vapour are in contact, and in equilibrium. Call this the steam line, or vaporization curve. The ordinate of *O* represents the vapour pressure of water at 0°; at lower temperatures the water freezes. Plot in a similar manner the vapour pressure of ice at different temperatures, and we get the curve *OP*, called the hoar-frost line, or the sublimation curve. Under these conditions, there is no intermediate liquid state, vapour condenses at once to a solid, and the solid passes directly into vapour. Solid iodine below its melting point 114° also vaporizes without liquefaction; arsenic can only be liquefied by melting the element under pressure; under ordinary conditions, arsenic sublimates without fusion. It is found that the effect of pressure on the melting point of ice can be represented by a curve *ON*, Fig. 61. The left-to-right downward slope of the curve shows that the melting point of ice¹ is lowered by increasing the pressure. Thus the melting point of ice at different pressures, according to W. Tammann (1900), is approximately:

Pressure	260	490	1100	1790	2020 atm.
Melting point	-2°	-4°	-10.11°	-17.6°	-20.59°

¹ At very high pressures, ice is said to form at least five different varieties, one of which is stable only at temperatures exceeding the melting point, 0° C., of ordinary ice.

and *in vacuo*, ice melts at $+0.0075^{\circ}$. To emphasize these relations, the curves in the diagram are slightly exaggerated. The curve *ON* is called the ice line or fusion curve.

Before progressing further in our study, it is convenient to fix special meanings to three terms; component, phase, and degree of freedom or variance.

Components.—The components of a system are those substances which take part in the reaction but which are not decomposed in the process. The components may be elements, or compounds which behave in a system, for the time being, as if they were elements. There is only one component in the system just considered, namely, water— H_2O ; the components in the system considered in the previous section—an aqueous solution of sodium chloride—are water (H_2O) and sodium chloride (NaCl); and two components were involved in our study of the solubility of sodium sulphate namely, water (H_2O) and sodium sulphate (Na_2SO_4).

Phases.—The components may group themselves in various ways. They may pass from one physical state to another, as when water boils or freezes; they may form simple solutions, as when salt dissolves in water; they may combine with one another in various ways, as when sodium sulphate (Na_2SO_4) forms the decahydrate ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$), etc. Every homogeneous state—solid, liquid, or gaseous—which the components can produce is called a phase. The phases of a system are the physical states in which the components can exist. A eutectic is not a phase. In the case of aqueous solutions of sodium chloride, the eutectic contained two phases— NaCl and H_2O . With an aqueous solution of sodium sulphate at the transition point, Fig. 58, we had to deal with four phases— Na_2SO_4 ; $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$; the saturated solution; and the vapour arising from the solution. With freezing water, we have the three phases: ice, water, and vapour. In homogeneous systems there can only be one phase, e.g. gaseous systems; and in heterogeneous systems there are always two or more phases.

Variance or degrees of freedom of a system.—It will be remembered that the condition of equilibrium of a gas with respect to temperature, pressure, and volume was defined (on p. 105) by the equation, $pv = RT$, for R is a numerical constant whose value depends upon the units of measurement (p. 97). If only one of these variables be fixed, say the volume, the state of the system will remain undefined, because the gas can retain one fixed volume, and yet have very different values for temperature and pressure. Two of the three variables must be known before the state of the system can be defined unequivocally, without ambiguity. If any two of the three variables be fixed, the third variable can only assume one definite value. The two fixed variables are said to be arbitrary or independent variables; the third variable which can be calculated from the condition of equilibrium ($pv = RT$) when the two independent variables are known, is called the dependent variable. Another term sometimes used for the independent variable is degree of freedom. The gaseous system under consideration has two degrees of freedom. The degree of freedom of a system—also called the variance¹ of a system—is the number of independent variables which must be fixed before the state of the system can be defined without ambiguity. The

¹ That is, "variableness," from the Latin, *variabilis*, variable.

gaseous system defined by the equation, $pv = RT$, is bivariant, because it has two independent variables, or two degrees of freedom.

The system consisting of liquid water and vapour has two variables: vapour pressure and temperature. So long as liquid water is present, the pressure is determined solely by its temperature; given either the pressure or the temperature, the other can be determined from the relation symbolized in the vapour pressure curve—Fig. 61. Hence the state of the system is defined by two variables—the one is dependent, the other independent. In other words, the system has one degree of freedom, that is, the system is univariant.

The three curves PO , OQ , and ON —Fig. 61—represent the conditions of equilibrium of three two-phase systems: solid-vapour, vapour-liquid, and solid-liquid respectively. These three curves meet at the point O . Here three phases can coexist in equilibrium. Hence the point O is called a triple point. The co-ordinates of the triple point are: pressure, 4.57 mm.; temperature, 0.0076°C . If the pressure or temperature be altered ever so little, one of the phases—ice or liquid water—will disappear and a two-phase univariant system represented by a point on one of the curves OP , OQ , ON will appear. At the triple point the system is invariant. Confining our attention, for the moment, to the liquid and solid, and neglecting the vapour, we can define: The freezing or melting point is the temperature at which both solid and liquid can exist side by side in contact with one another without changing.

Gibbs' phase rule.—J. W. Gibbs (1876-78) discovered an important relation between the number of components, the phases, and the degrees of freedom of a system. According to Gibbs' phase rule, a system will be in equilibrium when its variance is equal to the number of components in the system less the number of phases increased by 2. In symbols:

$$F = C - P + 2$$

where C denotes the number of components, P the number of phases, and F the variance or degrees of freedom of the system.

Invariant system.—An invariant system has no degrees of freedom, and the state of system cannot therefore survive a change of temperature or pressure. In that case $F = 0$, or $P = C + 2$. This means that the system will have $C + 2$ phases if it is in equilibrium. If there be one component in the system, as in the case of water at the triple point, three phases can coexist in equilibrium—ice, liquid water, and steam. Otherwise expressed, if a system has three phases and one component the phase rule tells us that it will be invariant, and therefore the slightest alteration of pressure or temperature will cause one of the phases to disappear.

Univariant systems.—These systems have one degree of freedom, and when the system is in equilibrium, $F = 1$, or $P = C + 1$. If one of the variables be known, the state of the system can be determined as indicated above.

Bivariant systems.—These systems have two degrees of freedom, and hence $F = 2$, or $P = C$. Two variables must be known before the state of the system can be determined. A saturated solution in the presence of an excess of the solute is univariant, but bivariant if not saturated. In the former case there are two components and three phases—solid, solution, and vapour; in the latter case there are two components and

two phases. Hence in the one case, $F = 2 + 2 - 3$; and in the other, $F = 2 + 2 - 2$. Again, in the region PON , Fig. 61, the system will be bivariant, because there is only one phase and one component. Pressure and temperature may be altered without interfering with the state of the aggregation of the ice so long as the variations keep within the boundary lines PO and ON . The same remarks may be applied to the condition of the water represented by points in the regions NOQ and POQ .

Object of the phase rule.—The phase rule is therefore a method of grouping systems which behave in a similar manner into one class. It is essentially a system for the classification of states of equilibrium. Systems having the same variance behave in an analogous manner under the influence of variations in temperature, pressure, and volume or concentration. It makes no difference whether the changes be chemical or physical. As indicated above, the phase rule also tells us whether the phases of a heterogeneous system are those necessary for equilibrium.

Special application to solids and liquids.—In the application of the phase rule to alloys, minerals, and solutions when the vapour pressure is negligibly small, only two variables need be considered—concentration or volume, and temperature. In that case, the phase rule reads:

$$F = C - P + 1$$

Granite, Fig. 3—composed of quartz, SiO_2 ; felspar, $\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$; and mica, say, $\text{K}_2\text{O} \cdot 3\text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$ —has three components: SiO_2 ; Al_2O_3 , and K_2O ; and three solid phases: mica, quartz, and felspar. The system is univariant. It is also in equilibrium, because, not being at a transition point, it is able to survive a small variation of temperature without changing the state of the system.

To summarize some of the systems previously considered:

TABLE VII.—CLASSIFICATION OF SYSTEMS BY THE PHASE RULE.

System.	Components.	Phases.	Degrees of freedom or variance.
Freezing water . .	Water	Liquid; solid; vapour	Invariant
Water above 0° . .	Water	Liquid; vapour	Univariant
Unsaturated solution of sodium chloride	Water, salt	Solution; vapour	Bivariant
Saturated solution of sodium sulphate at transition point	$\text{Na}_2\text{SO}_4; \text{H}_2\text{O}$	$\text{Na}_2\text{SO}_4; \text{Na}_2\text{SO}_4$ $10\text{H}_2\text{O}$; solution; vapour	Invariant
Freezing eutectic— sodium chloride and water . . .	Water, salt	Two solids; one liquid (neglect vapour)	Invariant
Steam and metallic iron in a closed vessel	Iron, oxygen, hydrogen	$\text{H}_2; \text{H}_2\text{O}; \text{Fe};$ Fe_3O_4	Univariant
Solution of oxygen in water	Oxygen; water	Gas; liquid	Bivariant

§ 8. Undercooling—Supersaturation.

A metastable system, though stable, is constantly menaced by a spontaneous transformation.—G. URBAIN.

Undercooling.—Heat a solution of sodium thiosulphate to, say, 70° in a glass flask; stir the molten mass with a thermometer as it cools; read the thermometer every two minutes, and finally plot the results on squared paper. A curve—called a **cooling curve**—resembling that illustrated in Fig. 62, *A*, will be obtained. The terrace in the cooling curve at 56° shows that a change of some kind takes place in the nature of the cooling substance at 56° . The terrace corresponds with the temperature at which solidification or freezing was in active progress. The sudden slackening in the rate of cooling corresponds with the “evolution of the latent heat of fusion” as the liquid solidifies. Repeat the experiment, but do not agitate the liquid; take care that the cooling liquid is quite still and protected from dust by, say, a loose plug of cotton wool in the neck of the flask. A cooling curve like that shown in Fig. 62, *B*, will be obtained. The liquid

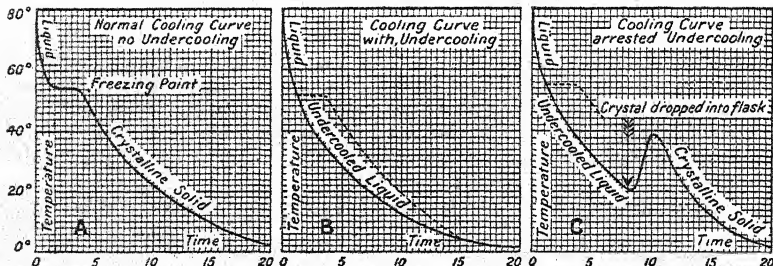


Fig. 62.—Cooling Curves of Molten Sodium Thiosulphate.

does not freeze, and no abnormal behaviour can be detected in the cooling curve. The liquid “ought to”¹ crystallize at 56° , but it does not. Drop a crystal of sodium thiosulphate into the liquid mass. The mass seems to solidify with almost explosive rapidity, and the thermometer immediately indicates a rise of temperature. The phenomenon is illustrated by Fig. 62, *C*. The liquid sodium thiosulphate at a temperature below 56° is said to be *surfused*, or, better, *undercooled*. The liquid may be kept in the surfused or undercooled condition an indefinite time, and the process of solidification can, in general, only be started by the introduction of a crystal of the same type as that which is formed during the solidification of the given substance. Often a fleck of the right kind of dust floating in the air suffices to upset the state of apparent equilibrium. Clear glasses and pottery glazes are solutions of silicates which have congealed to hard masses without crystallizing.

Supersaturation.—Similar phenomena occur if water be saturated with Glauber's salt— $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ —at 30° . Make sure that no excess of solid is in contact with the liquid, and let the solution cool as before—

¹ “Ought to” is put in inverted commas to show that the ordinary meaning of these words is modified.

without agitation and without dust. The solubility curve, Fig. 58, tells us that the solid "ought to" separate from the system as the temperature is reduced. Here is another case of apparent, false, or metastable equilibrium. Although the solution can be kept an indefinite time in this condition, the seeding, or inoculation, of a supersaturated solution by the introduction a very minute quantity of a crystal of dissolved salt will upset the state of apparent equilibrium.¹ The crystal fragment becomes the centre or nucleus from which crystals radiate into the solution on all sides. Similar results can be obtained with aqueous solutions of sodium acetate, sodium chlorate, etc. A supersaturated solution of ammonia alum, $\text{Al}_2(\text{SO}_4)_3 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 24\text{H}_2\text{O}$, deposits the same salt if seeded with crystal fragments of that salt; but A. Piccini and V. Fortini (1902) showed that if the saturated solution be sown with octahydrated ammonium thallic sulphate, the salt $\text{Al}_2(\text{SO}_4)_3 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 8\text{H}_2\text{O}$ appears. If a solution holds more salt than corresponds with the normal solubility curve of the salt, the solution is said to be supersaturated.

It is possible to distinguish between a saturated, an unsaturated, and a supersaturated solution by bringing each in contact with more of the solid. If the solution is unsaturated, more solid will dissolve; if saturated, none will dissolve; and if supersaturated, solid will separate until the solution is saturated. The concentration of an unsaturated solution is less, while the concentration of a supersaturated solution is greater than that of a saturated solution.

Related phenomena.—Many other examples of related phenomena are known. In analytical work the slow appearance of precipitates in dilute solutions is very common. Pure water may be easily cooled to -3° or -4° without the appearance of ice if kept quite still while the temperature is reduced; and the water can be easily cooled to -6° or -7° if a layer of oil be placed over the surface of the cooling water. *OR*, Fig. 61, represents the vapour pressure curve of the undercooled liquid. Phosphorus, sulphur, etc., behave in a similar manner. Ice has not been heated above 0° without melting, but liquid water can be heated to 105° or 106° without boiling. When the boiling does start, it proceeds with almost explosive violence. The phenomenon is called bumping. By suspending drops of water in a mixture of olive and linseed oils which has the same specific gravity as water, and a high boiling point, L. Dufour (1863) raised water to 178° without boiling. P. J. Coulier (1875) found that dust-free air saturated with moisture may be cooled below the normal temperature of condensation; and John Aitken (1880) showed that dust is necessary for the formation of fogs and raindrops, so that in perfectly clean dust-free air aqueous vapour does not condense, and mist does not form; without solid nuclei the cooling vapour becomes supersaturated.

Metastable and labile states of supersaturation.—Inoculation or seeding is usually necessary to start the process of crystallization of a supersaturated solution; and yet the supersaturation may be carried so far that the crystals will grow spontaneously in the solution, without

¹ In H. A. Miers' experiments on the crystallization of the rare organic compounds, salol and betol, it was found that the substances did not crystallize at first when allowed to cool in open vessels in the laboratory; but after a time, when the air of the laboratory had become impregnated with dust, presumably containing minute grains of both substances, crystallization readily occurred in open vessels exposed in the laboratory.

seeding. Indeed, it is possible to draw a "supersaturated solubility curve" representing the concentration of a solution at different temperatures where the supersaturation

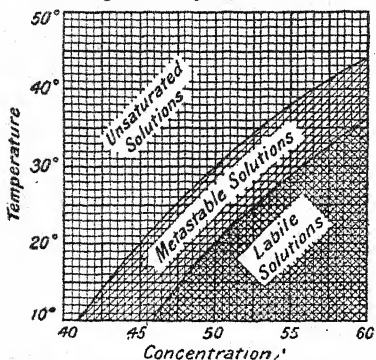


FIG. 63.—Labile and Metastable Equilibria of Saturated Solutions (after H. A. Miers).

it contains 45.8 per cent. of salt; it is in the labile state when it contains over 48.8 per cent.; and it is in the metastable state between these two concentrations.

Supersaturation and the phase rule.—The phase rule, it will be observed, applies to systems in real equilibrium, not to systems in a state of apparent, false, metastable equilibrium. We are repeatedly confronted with those little-understood phenomena which, for convenience, have been grouped under the general term "passive resistance," p. 136.

The kinetic theory of supersaturation.—The kinetic theory throws a little light on to the phenomenon of supersaturation. A saturated solution in contact with the solid is supposed to be closely analogous with a closed vessel containing a liquid in contact with its vapour. When in equilibrium, the same number of molecules pass from the surface of the solid into the solution and return from the liquid to the surface of the solid. If the state of equilibrium be disturbed by evaporation or by lowering the temperature, the equality of the two opposing actions is disturbed and a new condition of equilibrium is established. In the case of a metastable supersaturated solution, the exchange of molecules cannot take place because no free solid is present. Directly a particle of the dissolved solid is added, the dissolved substance is rapidly deposited about the submerged particle as a nucleus until the concentration of the solution has reached its normal value. In the case of a labile supersaturated solution, the crowding of the molecules is so great that they are able to form aggregates large enough to serve as nuclei for the separation of more solid.

Questions.

1. How would you proceed in order to find the volumes in which hydrogen and oxygen combine to form water? What modification of your apparatus would you make, if you were asked, in addition, to measure the volume of steam formed? What volume of gas is formed when 72 grams of water are decomposed

(a) electrolytically, (b) by means of sodium, (c) by heated iron ? [$H = 1$, $O = 16$].

—*Victoria Univ., Manchester.*

2. Write a short account of the physical properties of ice, water, and steam, comparing them with those of solids, liquids, and gases in general.—*London Univ.*

3. How would you prove that the composition of water may be expressed by the formula H_2O ? Mention everything that this formula implies. What volume would 9 grams of aqueous vapour occupy at $273^\circ C.$, and under 380 mm. pressure ? —*Aberdeen Univ.*

4. The specific gravity of a solution of hydrochloric acid is 1.175, and it contains 34.5 per cent. of the compound HCl . How many cubic centimetres of the solution will be required for the neutralization of 50 grms. of sodium hydroxide, $NaOH$, given the reaction : $HCl + NaOH = H_2O + NaCl$?

5. Give an account of the properties of water, especially those which may be termed chemical properties. How would you identify water and distinguish it from other colourless liquids ? —*London Univ.*

6. Write a short account of the phase rule, and indicate clearly some useful applications of this generalization.—*Board of Educ.*

7. When ordinary water which has been in contact with the air is boiled, gas is given off. How would you collect a sufficient quantity for analyses ? How could you prove that the gas was dissolved in water, and that it is not a product of the decomposition of water by heat ? How would you determine the amount of one of the constituents ? —*Oxford Locals.*

8. The formula for water has been written at different times HO and H_2O ($O = 8$). Point out the objections to both expressions and state concisely the reasons for representing the molecule of water by H_2O ($O = 16$).—*London Univ.*

9. Describe some case of chemical combination and some case of solution. Compare the two, and discuss the question whether any distinction should or should not be made between solution and other cases of chemical change.—*New Zealand Univ.*

10. Compare the bubbling of molten potassium chlorate at a high temperature with the bubbling of water when heated to about 100° .

11. 100 grams of water dissolve the following quantities of zinc sulphate at the temperatures named :

Temperature . . .	0°	25°	39°	50°	70°	80°	90°	100°
$ZnSO_4$	41.9	57.9	70.1	76.8	88.7	86.6	83.7	80.8 grams.

Plot the results on squared paper so as to show a "solubility curve."

12. Define the terms : degree of freedom, phase, component, and variant in Gibbs' rule. In the case of a system containing salt and water, and the phases salt, saturated solution, and vapour, how many variants must be fixed to establish equilibrium ? Discuss the reasons for your statement.—*Sydney Univ.*

13. Give an account of the phase rule and of the help it affords in understanding equilibria. Illustrate your answer in particular by applying the rule to the case of water.—*Bombay Univ.*

14. Can any distinction be drawn between the action of water on a soluble substance and other cases of chemical combination ? What are the resemblances, and what the differences between a substance in dilute solution and the same substance in the gaseous state ? —*New Zealand Univ.*

15. What are the chief characteristics of the solid, liquid, and gaseous states of matter ? What is the usual effect on each class of (a) alterations of temperature, and (b) alterations of pressure ? Describe experiments to illustrate your answer.—*Sydney Univ.*

16. Plot curves to represent the following tabulated numbers representing the number of grams of salt in 100 grams of water at the temperatures named :

	0°	10°	20°	30°	40°	50°	60°	70°	80°	90°	100°
Na_2SO_4 . . .	5	9	19.4	40	48.8	46.7	45.3	44.4	43.7	43.1	42.5
K-alum . . .	3.9	7.5	15.1	22.0	30.3	44.1	66.6	90.7	134.5	209.3	257.5

Answer by reference to the resulting diagram ; (a) equal volumes of two solutions on evaporation yielded 26.5 grams of sodium sulphate and potash-alum respectively. At what temperatures were the two solutions made up ? A saturated solution of potash-alum was made up at 64° , and another of sodium sulphate at 72° . If both solutions were cooled to 15° , how many more grams of potash-alum would separate out than of sodium sulphate ? —*Staffs County Major Scholarship.*

17. Some paraffin wax is heated in a test-tube until it melts and rises consider-

ably above its melting point in temperature. A thermometer is inserted in the melted wax and readings of the temperature are taken every minute until the wax has reached again the temperature of the room. The readings of the thermometer are now plotted against the corresponding times. Draw a diagram showing the sort of curve thus obtained, and explain its general character. How does it show the solidifying point of paraffin?—*Adelaide Pub. Exam. Board.*

18. What facts would you adduce to prove that the molecule of water contains two atoms of hydrogen and not more or less than two atoms?—*Science and Art Dept.*

19. How would you prove that the body formed by the combustion of hydrogen in air is water?—*Science and Art Dept.*

20. A writer says that water "can exist in three phases, namely, ice, water, and steam. These three phases together form a system in which there is one component, namely, the compound dihydrogen oxide (H_2O). Substituting the respective values for P and C , the equation becomes $3 + F = 1 + 2$. Hence, $F = 0$, so that under no condition can all three phases exist simultaneously in equilibrium." How has the writer erred in making this false conclusion?

CHAPTER X

CRYSTALS AND CRYSTALLIZATION

§ 1. The Crystallization of Salts from Solutions.

The world is not a meaningless medley. We do not believe that blind chance reigns supreme. On the contrary, we see order everywhere, and law is the regulating principle in all things and processes.—P. CARUS.

If a saturated solution of a salt be allowed to evaporate, crystals of the salt separate when the concentration of the solution becomes greater than that represented by a point on the solubility curve. The phenomenon becomes a little more complex when the solution contains two or more salts which do not act upon one another.

The solubility of a mixture of sodium and potassium chlorides in water at 25° is represented by the curves shown in Fig. 64. The ordinates represent quantities of sodium chloride; the abscissæ, quantities of potassium chloride. The concentration of a saturated solution of sodium chloride at 25° is represented by the point *A*, and of a saturated solution of potassium chloride by the point *B*. The line *AC* represents the composition of solutions of sodium chloride saturated in presence of the proportions of potassium chloride indicated by the abscissæ of the curve *AC*; and the line *CB*, the composition of solutions of potassium chloride saturated in presence of the proportions of sodium chloride represented by the ordinates of *CB*. The point *C* represents the composition of a solution saturated with both salts. The volume of solution is here understood to be the amount necessary to dissolve *A* and *B* gram molecules of the respective salts.

Sodium and potassium chlorides form neither "double salts" nor hydrates. Hence when a solution containing equal molecular proportions of both salts is evaporated, the solution will become saturated first with respect to the less soluble potassium chloride, and this salt will accordingly crystallize from the solution first. As evaporation proceeds, potassium chloride continues separating; the successive states of the

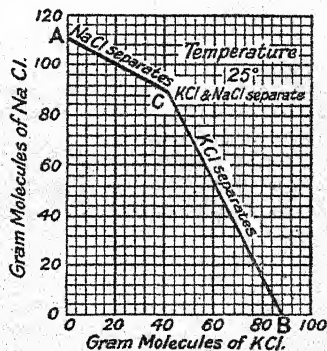


FIG. 64.—Solubility Curves of Mixed Solutions of Sodium and Potassium Chlorides.

solution are represented by points travelling in the direction *BC*. The phenomenon is really wonderful. The molecules of both sodium and potassium chlorides are uniformly diffused throughout the original solution; but, as soon as the evaporating liquid has attained a certain concentration, the molecules of the potassium chloride alone commence crystal-building; an ordered cosmos grows out of a chaotic mixture of molecules. The molecules of this salt withdraw from the solution in harmonious order, which is followed as rigorously as the bricklayer, when building a mansion, places brick upon brick, according to the plan predetermined by the architect. The operation continues until the solution has the composition represented by the point *C*. At *C* the solution is saturated with respect to both salts. Any further concentration of the solution will result in the deposition of sodium and potassium chlorides side by side, and the solution will go on depositing a mixture of the two salts until it has been evaporated to dryness. Here the molecules of both salts are simultaneously building crystals side by side. If sodium chloride be in excess, what has been said of potassium chloride along *BC*, applies to sodium chloride along *AC*. In each case *C* represents the end point of the crystallization.

§ 2. Fractional Crystallization.

What chemist who has watched under the microscope the beautiful symmetrical manner in which minute particles of a substance separating in the solid form from solution arrange themselves in geometrical figures, obeying well-established mathematical laws, can pretend to explain the cause of the astonishing behaviour of inert lifeless matter?—H. C. BOLTON

The molecules of a substance in solution appear to be distinct individuals before crystallization, while in a crystallizing solution each molecule appears to exert some specific attraction on its fellow molecules to enable them to separate from the solution in a definite orderly way. The chemical student sees the phenomena of crystallization every day, but he does not always realize the miracles being performed in his test-tubes. Some one has cited the formation of alum as an illustration. Here, the alignment of the molecules in a growing crystal can proceed so rapidly that a few seconds of our time must appear as a long era in the molecular world. Thus, if concentrated solutions of aluminium sulphate and potassium sulphate are mixed with constant stirring, a mass of beautiful transparent sparkling crystals of alum is immediately precipitated. The molecule of alum is usually represented $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$, and accordingly, in these few seconds, the atoms have had ample time to arrange themselves in molecular groups, each containing 96 atoms; and the molecules in turn have had time to align themselves in a precise methodical way to form an indefinite number of octahedral crystals containing more molecules than could be enumerated by continuous counting for myriads of years.

When a solution of two or more salts is slowly evaporated, if the salts differ appreciably in solubility one salt may pass more or less completely out of solution before the other commences to separate—fractional crystallization—provided at the temperature of separation the salts are independent of one another and exhibit no tendency to unite chemically or physically. Warm 50 grams of potassium chlorate, just above its

melting point, in a new porcelain dish, and keep the mass at that temperature until it becomes viscid and almost solid. This will occupy from ten to fifteen minutes. Let the mass cool. It contains undecomposed potassium chlorate, some potassium chloride, and potassium perchlorate.

Add 50 c.c. of hot water, say at 50°, and when all has disintegrated and the solution cooled, the crop of crystals of potassium perchlorate can be filtered off. Evaporate the filtrate until a drop crystallizes when rubbed on a cold surface. The first crop of crystals which separates as the solution cools is mainly potassium perchlorate, because this salt is so very much less soluble than the other two; 100 c.c. of water, at 15°, holds in solution about 36 grams of potassium chloride, 6.6 grams of the chlorate, and 1.5 grams of the perchlorate. The solubility curves of these three salts are shown in Fig. 65. If the evaporation be carried too far, crystals of

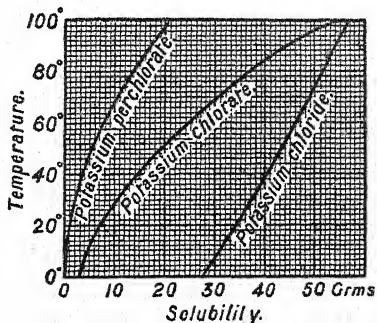


FIG. 65.—Solubility Curves of Potassium Chloride, Chlorate, and Perchlorate.

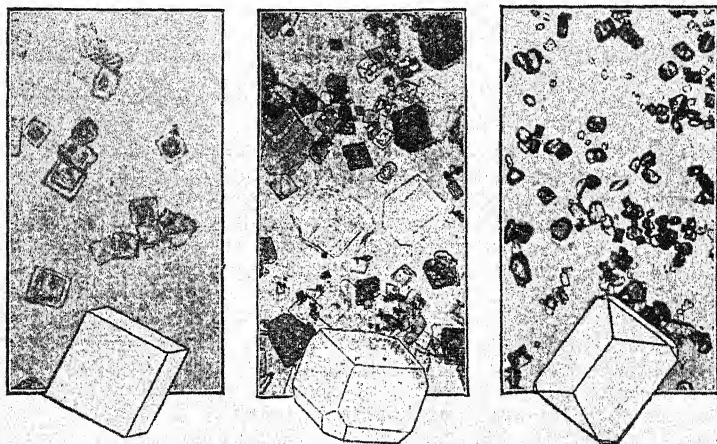


FIG. 66.—Crystals of Potassium Chloride, Potassium Chlorate, and Potassium Perchlorate.

potassium chlorate will separate. The first crop of crystals is redissolved in as small a quantity of hot water as is practicable, and again allowed to crystallize; potassium perchlorate can thus be obtained almost free from the other two salts; and by repeated recrystallization it is possible to isolate the salt in a high degree of purity. Recrystallization

is needed to get the pure product because the crystallizing salt often carries down with it some of the mother liquid, or some of the other salts dissolved in the mother liquid. If the evaporation be continued, potassium chlorate will eventually separate, and the process of recrystallization must be repeated, in order to isolate the pure salt. The remaining mother liquid is nearly all potassium chloride.

This operation—fractional crystallization—is sometimes a useful method of separating salts which differ appreciably in solubility. In some cases it is the only satisfactory method of separation, even though the salts in solution do not differ very much in solubility. The process of fractional crystallization is then very laborious, involving, may be, scores of crystallizations and recrystallizations. In other cases it is impossible to separate the salts in this way, because double salts separate. The microscopic appearance of the three salts isolated in this manner is shown in Fig. 66. The outline drawings represent perfect crystals. As a rule, the slower the process of crystallization, the larger and more perfect the crystals. The crystals are usually more or less distorted because, owing to local differences in the concentration of the crystallizing solution, the solution in the vicinity of some faces of the crystal may be more concentrated than the others. Perfect crystals are rarely found in nature or in the laboratory.

§ 3. Crystals.

In whatever manner, or under whatever circumstances, a crystal may have been formed, whether in the laboratory of the chemist or in the workshop of nature, in the bodies of animals or in the tissues of plants, up in the sky or in the depths of the earth, whether so rapidly that we may literally see its growth, or by the slow aggregation of its molecules during perhaps hundreds, perhaps thousands of years, we always find that the arrangement of the faces of the crystal, and therefore its other physical properties, are subject to fixed and definite laws.—H. P. GURNEY.

The constancy of interfacial angles.—When homogeneous substances solidify from a state of vapour, fusion, or solution, their particles cohere so as to form solid figures—crystals—with regular symmetrical shapes bounded by plane faces. The solids are then said to be crystallized. Every crystalline substance of definite chemical composition has a specific crystalline form characteristic of that substance. This is sometimes called R. J. Haüy's law. The faces of crystals of the same substance may vary in size and shape; but if the crystals possess the same chemical composition, and are at the same temperature, the interfacial angles have the same numerical value. In other words, the angles between similar faces of crystals of the same substance are precisely the same, and are characteristic of that substance. This "law" was first announced by D. Guglielmimi (1688). This means that the crystalline form of a substance is not determined by the absolute position nor by the sizes of the faces of the crystal, but rather by the dimensions of the interfacial angles. The primary dominant faces, so to speak, may persist while the angles and edges of some of the crystals may be truncated and bevelled, giving rise to new facets. In spite of these variations, the crystals of a given substance always retain its fundamental form. Different substances may have a similar chemical composition and different interfacial angles.

Polymorphism.—Crystals of different substances have different forms; crystals of the same substance developed under the same conditions have the same form; but crystals of the same substance developed under different conditions may or may not have the same form. For instance, crystals of sulphur formed above or below 94.5° , p. 475; there are two differently shaped crystals of sodium phosphate, Fig. 67; crystals of sodium chloride are octahedral if grown in alkaline solutions, and cubical in neutral solutions; and conversely, crystals of alum are usually octahedral, but cubical if grown in alkaline solutions. Substances which crystallize in two different forms are said to be dimorphous, Fig. 67; and substances which crystallize in three different forms are said to be trimorphous. Titanic oxide, TiO_2 , for example, is known in three forms, rutile, anatase, and brookite. *Polymorphism* is the general term applied to the phenomenon when a substance crystallizes in more than one form.

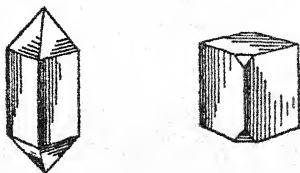


FIG. 67.—Dimorphic Forms of Sodium Biphosphate.

Ideal and distorted crystals.—One or more faces of a crystal may be abnormally developed or stunted in growth. During the growth of a crystal, the concentration of the mother liquid is rarely so evenly balanced on all sides as to allow the growth to proceed with the same rapidity in all directions. The crystal will grow fastest where the solution is most concentrated. If a crystal grows on the bottom of a liquid at rest, flat plates, almost parallel with the bottom of the vessel, may be formed; while if the solution be agitated during crystallization, a more uniform growth in all directions may prevail. This is not all, a cubical crystal may habitually grow most rapidly in *one* direction so as to form a needle-like, acicular, or prismatic crystal, *A*, Fig. 68; or in *two* directions so as to form a tabular or plate-like crystal, *B*, Fig. 68; or the crystal may grow uniformly in all *three* directions and the ideal cube, *C*, Fig. 68, will be produced. The angles between the faces, however, will remain unchanged: 90° . The most suitable condition for uniform growth in all directions occurs when the growing crystal is suspended in the middle of the given solution by means of a thread. The crystal then approximates more or less closely to the ideal form. If the crystallization of a solution be rapid, the crystals are usually much smaller than if the process be slow. Again, the habit of a crystal may change when grown from liquids containing other salts in solution. Thus, potassium chlorate gives prismatic crystals (Fig. 66) when grown in aqueous solutions; if much calcium chloride be present, small needle-like crystals are obtained; and if potassium iodate be present in the solution, long plate-like crystals separate from the solution. The crystal angles, however, are the same in each case. The general shape of a crystal or the crystal habit may vary in that the different sets of planes which bound the crystal may grow unequally, and some may even be

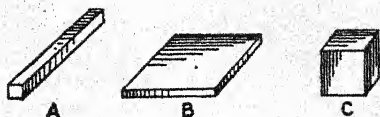


FIG. 68.—Ideal and Distorted Crystals.

suppressed. Thus, calcite occurs as rhombohedral, scalenohedral, and hexagonal prisms in almost infinite variety, but all can be referred to one primary or fundamental form.

Ideal crystals.—However much the crystals may be distorted, the angles between like faces have the same value, and in consequence, similar faces can be imagined at the same distance from the centre of the crystal. In this way, the ideal form of a crystal can be derived from that of a distorted crystal.

§ 4. The Seven Styles of Crystal Architecture.

A symmetrical shape is one which consists of parts exactly similar, repeated a certain number of times, and placed so as to correspond with each other. The symmetrical parts of a crystal are, under like circumstances, alike affected.—W. WHEWELL.

Planes of symmetry.—The great variety of crystals formed by different substances can be arranged into seven systems. Certain qualities are peculiar to the members of each system. For instance, the planes of symmetry of each class are characteristic. A plane of symmetry is an imaginary plane which divides the crystal into two parts such that one part is the exact but inverse counterpart of the other. In other words, the two parts bear to one another the same relation that the image in a mirror bears to its object. The mirror is the equivalent of a plane of symmetry. A crystal of sodium chloride, for example, has nine planes of

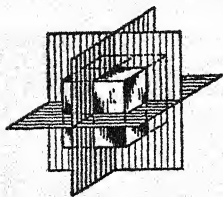


FIG. 69.—Sodium Chloride.



Planes of Symmetry.

FIG. 70.—Gypsum.

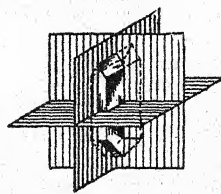


FIG. 71.—Zinc Sulphate.

symmetry; three are indicated in Fig. 69, and six others are obtained by taking planes diagonally through three faces of the cube. The crystal of gypsum, Fig. 70, has only one plane of symmetry; and a crystal of zinc sulphate has three planes of symmetry, Fig. 71.¹

Axes of symmetry.—Then again, a crystal may be rotated about a definite axis through an angle, which is a simple fraction— $\frac{1}{2}$, $\frac{1}{3}$, $\frac{1}{4}$, or $\frac{1}{6}$ of the angle of complete rotation, 360° —such that the faces, edges, and corners are brought into similar or symmetrical positions, and the aspect of the crystal is the same as before rotation. The axes of rotation are then called axes of symmetry. Thus we speak of dyad, triad, tetrad, and hexad axes of symmetry according as there are 2, 3, 4, or 6 positions of symmetry during a complete rotation. Thus Fig. 72, A, represents a horizontal cross-section of a crystal with one hexad axis of symmetry,

¹ Compare also Fig. 5 for zinc sulphate.

because during the rotation of the crystal about the axis O , there are six positions where the original aspect of the crystal is the same. Fig. 72, B, C, D, respectively, denote tetrad, triad, and dyad axes of symmetry. Other grades of symmetry, pentad, heptad, octad, etc., are not possible in crystals. The study of the forms of crystals is a special branch of chemical physics—*crystallography*. Sufficient is here stated to give the student an idea of a few special terms in common use.

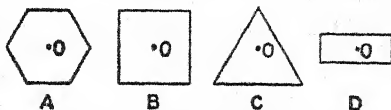


FIG. 72.—Axes of Symmetry.

The seven systems of crystal architecture.—Every known crystal can be referred to one of the following seven systems:

I. Triclinic system.—Crystals of this system have no axes nor planes of symmetry. This system has also been designated the “anorthic,” “clino-rhomboidal,” “asymmetric,” or the “double oblique” system.

EXAMPLES.—Potassium dichromate; copper sulphate— $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$; calcium thiosulphate— $\text{CaS}_2\text{O}_3 \cdot 6\text{H}_2\text{O}$; boric acid; potassium ferricyanide; anhydrous manganese sulphate; copper selenate; anorthite—lime felspar; cryolite; etc.

II. Monoclinic system.—Members of this system have one plane of symmetry, or one dyad axis of symmetry, or both a plane and dyad axis. This system has also been styled the “monosymmetric,” “clinorhombic,” or the “oblique” system.

EXAMPLES.—Borax— $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$; gypsum; ferrous sulphate— $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$; sodium carbonate— $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$; felspar—orthoclase; sodium sulphate— $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$; ammonium magnesium sulphate— $\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 6\text{H}_2\text{O}$; potassium chlorate; potassium tetrathionate— $\text{K}_2\text{S}_4\text{O}_6$; tartaric acid; sulphur—from fusion; cane sugar; arsenic disulphide—realgar; etc.

III. Rhombic system.—Here the crystals may have three planes of symmetry, and three dyad axes of symmetry. This system is sometimes called the “orthorhombic,” “trimetric,” or the “prismatic” system.

EXAMPLES.—Zinc sulphate— $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$; magnesium sulphate— $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$; ammonium magnesium phosphate— $\text{NH}_4\text{MgPO}_4 \cdot 6\text{H}_2\text{O}$; potassium sulphate; arragonite; anhydrous sodium or silver sulphate; sulphur from solution; barytes; sodium arsenate; sodium phosphate— $\text{NaH}_2\text{PO}_4 \cdot \text{H}_2\text{O}$; iodine; potassium nitrate; tartar emetic; potassium chlorate; potassium permanganate; topaz; marcasite; pyrrhotite; tin; tridymite; silver nitrate; lead carbonate; silver sulphide; ferrous sulphate— $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$, etc.

IV. Tetragonal system.—The members of this system may have five planes or symmetry, one tetrad, and maybe four dyad axes of symmetry. This system is sometimes called the “pyramidal,” “quadratic,” or the “quaternary” system.

EXAMPLES.—Rutile; cassiterite; zircon; mercurous chloride; potassium ferrocyanide; nickel sulphate; potassium hydrogen phosphate— KH_2PO_4 ; native lead molybdate— PbMoO_4 ; sodium meta-antimonite— NaSbO_3 ; potassium hydrogen arsenate— KH_2AsO_4 ; scheelite; tin; strychnine sulphate; etc.

V. Trigonal system.—The crystals of this system may have three planes of symmetry, one triad, and three dyad axes of symmetry. This system is sometimes called the “rhombohedral” system, and it is sometimes regarded as a special development of the hexagonal system.

EXAMPLES.—Sodium periodate— $\text{NaIO}_4 \cdot 3\text{H}_2\text{O}$; quartz; tourmaline; antimony; bismuth; calcite; ice; graphite; sodium nitrate; arsenic; nickel

sulphide—millerite; cinnabar; calcium chloride— $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$; corundum; cadmium carbonate; bismuth iodide; ferrous carbonate; zinc carbonate; manganese carbonate; etc.

VI. Hexagonal system.—Here the crystals may have seven planes of symmetry, one hexad, and six dyad axes of symmetry.

EXAMPLES.—Beryl; apatite; cadmium sulphite—greenockite; copper sulphide; lead iodide; magnesium; beryllium; zinc; cadmium; calcium; pyrrhotite; etc.

VII. Cubic system.—The crystals in this system may have nine planes of symmetry, six dyad, three tetrad, and four triad axes of symmetry. This system has been variously styled the "isometric," "regular," "tesseral," "octahedral," or the "tessural" system.

EXAMPLES.—Diamond; potassium chloride; sodium chloride; alum; fluor-spar; iron pyrites; lead nitrate; magnetic oxide of iron; barium nitrate; arsenic trioxide; galena; garnet; ammonium chloroplatinate; silver chloride; boracite; iron; platinum; lead; phosphorus; gold; copper; silver; arsenic; etc.

§ 5. The Internal Structure of Crystals.

Imagine two hundred brilliant violin players playing the same piece with perfectly tuned instruments, but commencing at different places selected at random. The effect would not be pleasing, and even the finest ear could not recognize what was being played. Such music is made for us by the molecules of gases, liquids, and ordinary solids. They may be highly gifted molecules with a marvellous internal structure, but in their activity, each disturbs the others. A crystal, on the other hand, corresponds with the orchestra led by a vigorous conductor when all eyes intently follow his nod, and all hands follow the exact beat. This picture enables us to understand how crystals can exhibit whole ranges of phenomena quite wanting in other bodies. To me, the music of physical law sounds forth in no other department in such full and rich accord as in crystal physics.—W. VOIGT.

Crystals are not only peculiar in the regularity of their external shape, but they also possess a definite internal structure. The properties of crystals are not always the same in different directions. The hardness, elasticity, crushing strength, rate of solution in acids, optical, thermal, and electrical properties are generally different in different directions. This means that the elasticity, refraction of light, thermal expansion, etc., of a crystal is usually different when measured in different directions.

In illustration, let a slice of quartz be cut perpendicular to the long axis, and another slice be cut perpendicular to this; let each slice be covered with wax, and pierced at the centre so that a wire can be inserted and heated by an electric current. The wax naturally melts about the hot wire. In the former case, the molten wax will form a circle; and in the latter case, an ellipse. If a crystal of calcite be hung in a beaker of dilute hydrochloric acid by means of a platinum wire, solution does not occur at a uniform rate over the whole surface, but the crystal dissolves more rapidly in one direction than in another. With salt, cubic system, the rate of solution is the same in all directions. Again, the rate at which light travels through cubic crystals is the same in all directions, but not with members of the other systems.

The external form of crystals is their most obtrusive characteristic, and it was naturally the first to arrest attention; but the geometrical shape is by no means the most characteristic property of crystals, because

the external geometrical form may be destroyed, and yet the fragments do not cease to be crystals. On the contrary, the most perfect glass model of a crystal is not a crystal, because it lacks the characteristic internal properties of crystals.¹ In the case of granite (Fig. 3), the crystals of feldspar, quartz, and mica have been so crowded during their growth that they have had no chance to develop their characteristic external shape. The internal structure of each mineral, however, is characteristic. A crystal has therefore been defined as "a solid body bounded by plane surfaces arranged according to definite laws, and possessed of definite physical properties. Both the external form and the physical properties result from a definite internal structure." The essential difference between crystalline and amorphous substances is one of internal structure, not necessarily external shape.

An amorphous substance is one which, during "solidification," has not taken the definite external shape characteristic of crystals, and the properties when measured in any one direction are the same as when measured in any other direction. In this case it is assumed that the constituent molecules are arranged haphazard. In crystals, on the contrary, where the properties along parallel directions are the same, but different in directions that are not parallel, it is assumed that the ultimate molecules, or their motions, are oriented or arranged in a definite regular manner. W. Voigt (1906) aptly illustrates this idea by the metaphor cited above.

§ 6. The Growth of Crystals.

The very molecules appear inspired with a desire for union and growth.—
J. TYNDALL.

We do not understand the phenomenon of crystallization, nor do we know how crystals grow. The facts indicated in the preceding sections have made us almost certain that crystals grow by accretion, molecule by molecule, like bricks in the hands of the builder, and in accord with "an architectural plan more elaborate and exact than that of any human architect." Whatever be the actual size and shape of the structural unit, it is equally certain that (1) *the structural units or molecules of crystals of the same chemical substance, under similar conditions, must be alike in size, and in the distribution of their attractive forces*; (2) *the relative position of any one molecule must be symmetrical with that of every other molecule*; and (3) *that the way the molecules are packed in and along all parallel planes must be the same*. No other arrangements can be regarded as possible in a crystal. Hence the study of crystal structure is reduced to the investigation of the possible arrangements of networks of structural units in space—called space lattices—which satisfy these conditions. Mathematicians have shown that 230 different methods of packing are possible, and that all these can be classified into seven groups corresponding with the seven systems of crystal architecture indicated above.

¹ The shapes of gems cut and polished to accentuate the ornamental value of the gem must not be confounded with crystal structure. Similarly, the term "crystal" applied to cut glass has a different meaning from the special use of the word "crystal" in the text. Transparent glass is not crystalline; some varieties of opaque glass are more or less crystalline.

It is shown in text-books on physics that when a beam of light strikes a series of very fine lines ruled regularly on the surface of a metal or glass plate, each line acts as a fresh centre from which a secondary train of light waves is diffracted. The diffracted rays enhance some of the normal light waves and damp down others; and the net result is that the beam of light is analysed into a series of spectra. A crystal behaves as a solid with a continuous surface when exposed to ordinary light, but M. Laue showed, in 1912, that with X-rays having a wave-length of the order 10^{-9} cm., the crystal acts as a kind of grating, owing to the regular disposition of the structural units. W. Friedrich and P. Knipping applied the idea and allowed the diffracted X-rays to impinge on a photographic plate after passage through the crystal. The diffracted rays produced spots on the developed paper which corresponded with structural units encountered by the rays in passing through the crystal. The idea was developed by W. H. Bragg and others, and the data so obtained allow the disposition of the structural units of crystals to be determined. As a result, the space lattices of the crystals of most of the elements and many of their compounds have been determined, and the distances apart of the different units measured. The diamond has interpenetrating cubic lattices, so also have silicon, germanium, and grey tin; while white tin



FIG. 73.—Body-centred Cubic Lattice.



FIG. 74.—Face-centred Cubic Lattice.

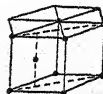


FIG. 75.—Close-packed Hexagonal Lattice.

has a tetragonal lattice. The elements Li, Na, K, Pb, V, Ta, Cr, Mo, W, and α -Fe have body-centred cubic lattices, Fig. 73; Cu, Ag, Au, Ca, Al, Pb, γ -Fe, Ni, Rh, Pd, Ir, and Pt have face-centred cubic lattices; Be, Mg, Zn, Cd, Hg, Ti, Zr, Ce, Co, Ru, and Os have close-packed hexagonal lattices, Fig. 75. The crystal unit of compounds may contain two or more molecules of the same substance. There is no sign of chemical combination because the packing of the units in the crystal are dominated by the crystalline cohesive forces, while the chemical forces binding the atoms in the molecule appear to play a subordinate part.

When a crystal is growing, curiously enough, the liquid in the immediate vicinity of the growing face is more concentrated, for it contains more of the dissolved substance per unit volume than the liquid a short distance away from the growing face. At first sight, it seems as if the growing crystal exerts some kind of attraction on the molecules of the dissolved substance a short distance away. For instance, if a saturated solution of zinc silicate in molten lead borosilicate tinted with cobalt silicate be allowed to crystallize, the crystals of willemite which separate will also abstract the cobalt silicate from the solution, and form patches of beautiful "azure blue" crystals in a colourless matrix. If no crystallization occurs, the matrix will be uniformly coloured an intense blue.

The kinetic theory of crystal growth.—The kinetic theory has taught us that, during crystallization, it is probable that a series of

exchanges between the molecules of the crystal and the molecules of the solution are going on all over the surface of the growing crystal. Molecules of the dissolved substance are attracted to the surface of the growing crystal, the molecules of the crystal continually pass into solution again. If the crystal is growing, more molecules are deposited on the crystal than are lost in unit time; and if the crystal is dissolving, less molecules are deposited on the crystal than are lost in a unit of time.

Let Fig. 76 represent, diagrammatically, a growing crystal, one face of which is incomplete; and assume—as R. Hooke did in 1665—that the structural units are spherical molecules. If a sphere lodges against a completed face, it can touch *three* other spheres, and whether or not the molecule leaves the growing crystal will depend upon the force of attraction exerted upon it by the three contiguous molecules. Again, suppose that a sphere lodges on the little ledge formed by the top layer of the incomplete face. It will then touch *five* instead of *three* spheres; and it will be held in place by the attraction of five contiguous spheres.

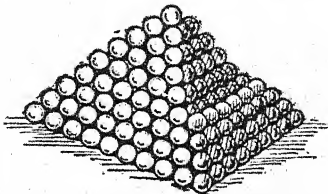


FIG. 76.—Imaginary Diagram of Growing Crystal.

Obviously, therefore, (1) during the exchange of molecules between the growing crystal and the solution, those molecules which have been deposited on the growing face will be retained more tenaciously than those deposited on a completed face; (2) as soon as a few molecules happen to be deposited in juxtaposition on the face of a crystal, subsequent growth on that face will be more rapid than the sporadic growth elsewhere; (3) an incompleting layer will rapidly extend until it covers the entire face of the crystal, etc. These deductions are in harmony with known facts. In illustration, if one part of a crystal be damaged the injured part will grow more rapidly than the other parts of a crystal until the injury disappears, and the perfect crystal is restored. Further, if a crystal be removed from a solution in which it is growing, it does

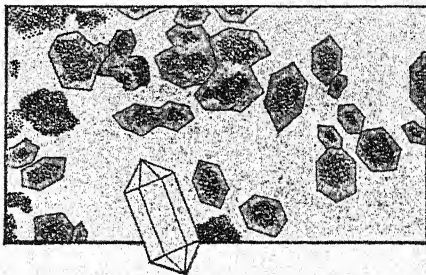


FIG. 77.—Growth of Quartz Crystals about Sand Grains.

not lose its power of growth, for if the crystal be placed in a suitable environment at any future time, it will continue growing as if there had been no interruption. These two statements are demonstrated by the sketch, Fig. 77, from a remarkable photograph. Some quartz crystals, grown during some former geological period, have lost their external crystalline form by attrition as they "knocked about the world"—blown about as sand in the deserts, washed down the hillside in streams of water, etc.—and they were finally deposited as rounded sand grains along with the mountain limestone

from some prehistoric sea. There the damaged crystals—sand grains—met a suitable environment in later years—probably water percolating through the limestone rocks, and carrying silicic acid in solution. The damaged crystals were repaired. Each sand grain, now imbedded in each of the repaired crystals, served as a foundation for rebuilding the damaged quartz crystals on the original architectural plan.

§ 7. Isomorphism—Mitscherlich's Law.

The laws of nature represent design; they are embodied design.—P. CARUS.

While making preparations of the arsenates and phosphates of potassium and ammonium, E. Mitscherlich, 1818, noticed that the crystals were so like each other as to be indistinguishable by simple inspection; and a closer examination led Mitscherlich to conclude: (1) that bodies of different chemical composition may have the same crystalline form; (2) substances of similar constitution have the same crystalline form. Mitscherlich wrote: "*The same number of atoms combined in the same manner produce the same crystalline form; the crystalline form is independent of the chemical nature of the atoms, and is determined solely by their number and mode of combination.*"—*Mitscherlich's law*. He noticed that the acid arsenates and phosphates of potassium and ammonium crystallize in similar tetragonal forms, and that one element or group of elements may be exchanged for another which appears to act in an analogous manner. Thus arsenic may be exchanged for phosphorus; and potassium for ammonium without affecting the form of the crystals. In Mitscherlich's words: "Every arsenate has its corresponding phosphate, composed according to the same proportions, combined with the same amount of water of crystallization, and endowed with the same physical properties; in fact, the two series of salts differ in no respect except that the radicle of the acid in the one series is phosphorus, while in the other it is arsenic."

Besides the phosphates and arsenates, Mitscherlich observed that mineral carbonates: calcite— CaCO_3 ; dolomites— $\text{CaMg}(\text{CO}_3)_2$; chalybite— FeCO_3 ; and dialogite— MnCO_3 —all form isomorphous crystals in the trigonal system (Fig. 137); again, the mineral sulphates: barytes— BaSO_4 ; celestine— SrSO_4 ; and anglesite— PbSO_4 —all form similar rhombic crystals; aragonite, CaCO_3 ; witherite, BaCO_3 ; strontianite— SrCO_3 ; and cerussite— PbCO_3 , form isomorphous rhombic crystals (Fig. 137); etc. Mitscherlich applied the term isomorphism—from the Greek *isos* (isos), equal; *μορφή* (morphe), shape—to connote the fact that analogous elements can replace one another without affecting the apparent shape of the crystals.

Mitscherlich's law of isomorphism, as well as the phenomena of polymorphism, appear to contradict Haüy's law (p. 206), but later investigations have shown that the crystals of isomorphous substances are not absolutely identical, but only similar in form. There are small but real differences between the members of an isomorphous series of compounds. Thus, A. E. H. Tutton found that in the isomorphous selenates and sulphates of potassium, rubidium, and caesium, specific chemical replacements are accompanied by clearly defined changes in the crystal structure along specific directions. Thus, when the basic element, say, potassium, in an

alkaline, sulphate or selenate is replaced by another of the same alkali family group, rubidium or caesium, the greatest alteration occurs in the crystal angles corresponding with an elongation of the vertical axis; and when the acid-forming element sulphur is replaced by selenium, its family analogue, the greatest expansion takes place along the horizontal axes of the crystals. The diagram, Fig. 78, shows, in an exaggerated manner, the effect of replacing potassium in potassium sulphate or selenate by the basic elements rubidium or caesium.

Retgers' law of mixed crystals.—Extended observations have multiplied examples of substances which possess a similar chemical constitution and a similar crystalline form; but at the same time the observations have also brought into prominence the fact that substances which crystallize in similar or identical forms—particularly in the cubic system—may exhibit wide divergencies in chemical constitution. The converse of Mitscherlich's law does not, therefore, hold good. Similarity of chemical composition or similarity in crystalline form are not adequate tests for isomorphism. Mitscherlich also stated that "while substances of different crystalline

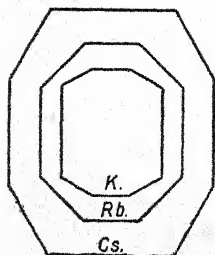


FIG. 78. (After A. E. H. Tutton.)

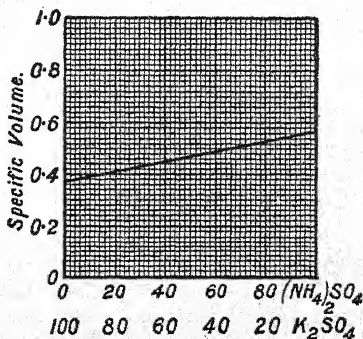


FIG. 79.—Specific Volumes of Mixed Crystals of Ammonium and Potassium Sulphates.

form cannot combine other than in fixed proportions, substances of the same crystalline form can crystallize together in all proportions"; and J. W. Retgers (1889) considers all important the property of forming mixed crystals in all proportions such that "if the percentages of one constituent of the mixture be plotted as abscissæ, and the corresponding magnitudes of the physical properties be plotted as ordinates, the different points lie in a continuous line," or the physical properties of isomorphous mixtures are continuous functions of the percentage composition—Retgers' law. Physical properties here include geometrical, optical, thermal, elastic, and electrical properties, but on practical grounds, Retgers considers specific gravity, or the reciprocal of specific gravity—specific volume—to be the most suitable property for investigation. An example is indicated in Fig. 79, where the specific volume of mixed crystals of potassium and ammonium sulphates are plotted. The continuity of the curve shows that the specific gravity or specific volume and chemical composition of the mixed crystals are isomorphous. The curves sometimes show a break, as is the case with Mitscherlich's phosphates, showing

that the two substances are not miscible in all proportions, but the two portions of the curve are parts of one straight line corresponding with the isomorphous character of the two salts. If the curve shows a kink or abrupt bend, the two salts, even if perfectly miscible in all proportions, would not, according to Retgers' definition, be called isomorphous. For instance, ammonium and ferric chlorides are not isomorphous although octahedral ammonium chloride forms coloured mixed crystals by taking up a small amount of ferric chloride. In special cases, double compounds may be formed which interfere with the application of Retgers' rule—e.g. potassium and silver nitrates form $\text{KAg}(\text{NO}_3)_2$; and potassium chloride and cupric chloride— $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$ —form $2\text{KCl} \cdot \text{CuCl}_2 \cdot 2\text{H}_2\text{O}$.

Overgrowths.—If a crystal of dark violet chromium alum be placed in a saturated solution of ordinary potassium alum, a transparent colourless overgrowth of potassium alum is deposited as a crust over the dark-coloured chromium alum as a nucleus. Similarly, a crystal of colourless zinc sulphate— $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ —can be coated with an overgrowth of green nickel sulphate— $\text{NiSO}_4 \cdot 7\text{H}_2\text{O}$; crystals of sodium nitrate grow on Iceland spar; and a pale amethyst triclinic crystal of manganese sulphate— $\text{MnSO}_4 \cdot 5\text{H}_2\text{O}$ —can be coated with blue copper sulphate— $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, etc. H. Kopp (1879) stated that this power of forming overgrowths, as well as the power of forming mixed crystals, enables isomorphism to be detected even when no particulars about the crystalline form or about the chemical composition are available. There are some exceptions to the test for isomorphism—trigonal potassium sulphate can be coated with a layer of hexagonal potassium sodium sulphate— KNaSO_4 .

Tests for isomorphism.—Substances which satisfy the three tests:

- (1) Similarity of crystalline form (Mitscherlich);
- (2) Formation of mixed crystals (Retgers); and
- (3) Formation of overgrowths (Kopp),

are generally of analogous chemical constitution, and accordingly isomorphous. None of these tests is an infallible criterion, and here, as is so often the case, a conclusion can be drawn only after carefully balancing the available circumstantial evidence.

Isodimorphism.—Each of the sulphates $\text{RSO}_4 \cdot 7\text{H}_2\text{O}$ (where R may be Mg, Zn, Ni, Co, Fe, Mn) is dimorphous, forming rhombic and also monoclinic crystals. The rhombic crystals of all the salts form one isomorphous series, and the monoclinic crystals of all the salts form another isomorphous series. This phenomenon—two independent series of isomorphous salts—is called isodimorphism. Monoclinic feldspar—orthoclase—often contains sodium; while triclinic soda feldspar—albite—often contains potassium. Hence P. Groth (1874) inferred that this is a case of isodimorphism, and that two pure varieties—monoclinic and triclinic soda and potash feldspars should exist. The prediction was verified two years later by the discovery of microcline, the triclinic form of potash feldspar, and later still by the discovery of barbierite, the monoclinic form of soda feldspar. Sodium phosphate forms two distinct crystals—rhombic and monoclinic. The arsenate appears in only one of these forms. Hence it is inferred that a monoclinic sodium arsenate isomorphous with rhombic sodium phosphate remains to be discovered.

Homeomorphism or isogonism.—Substances which differ in chemical constitution but have a similar crystalline form are said to be isogonic or

homeomorphic, to distinguish them from isomorphous substances. For example, aragonite, CaCO_3 ; nitre, KNO_3 ; barytes, BaSO_4 ; and potassium permanganate, KMnO_4 , have a similar crystalline form, but a different chemical constitution. The application of the X-ray diffraction spectrum enabled W. H. Bragg to show that the structures of crystals of calcium carbonate and potassium nitrate are similar. Again, cinnabar, HgS ; susannite, $\text{PbSO}_4 \cdot 3\text{PbCO}_3$; potassium hydrosulphate, KHSO_4 , and felspar, KAlSi_3O_8 , are not isomorphous, in spite of their external resemblance.

§ 8. The Rectification of Atomic Weights by Isomorphism.

The law of isomorphism can be used as a control in deducing the chemical composition of a salt; and also in atomic weight determinations for deciding between two numbers which are multiples of a common factor. The method is restricted to crystalline compounds; and it is only applicable in conjunction with other methods of atomic weight determinations since at least one member of the isomorphous series must be known.

Mitscherlich deduced the number 79 for the atomic weight of selenium by this method, and he also gave selenious and selenic acids formulæ corresponding with sulphurous and sulphuric acids respectively, on account of the isomorphism of the sulphates and the selenates. The analyses of potassium sulphate and of potassium selenate gave:

	Potassium.	Oxygen.	Sulphur.	Selenium.	Total.
Potassium sulphate	44.83	36.78	18.39	—	100.00
Potassium selenate	44.83	36.78	—	45.40	127.01

Assuming that the molecule of potassium sulphate contains one atom of sulphur; that the molecule of potassium selenate contains the same number of atoms; and that the atomic weight of sulphur is 32, we have:

$$\text{Atomic weight S : Atomic weight Se} = 18.39 : 45.34$$

Hence,

$$32 : \text{Atomic weight Se} = 18.39 : 45.34$$

$$\therefore \text{Atomic weight of selenium} = 79.00$$

About 1836 the atomic weight of copper was supposed to be 63.4, and of silver, 216.6. The analysis of the native sulphides of these elements were accordingly represented by the formulæ Cu_2S and AgS . But J. B. Dumas (1837) pointed out that the two minerals are isomorphous, and various mixed sulphides of the two elements are known by the general term: "fahlore." Hence the constitution of the two sulphides is probably the same. Assuming the formula of the one is Cu_2S , the formula of the other will probably be Ag_2S ; and the atomic weight of silver 108.3, not 216.6. This result agrees with evidence deduced from other independent sources. More exact determinations of the atomic weight of silver make this element 107.9; but this does not affect the principle of the argument.

EXAMPLE.—Analyses of alumina show that $\text{Al} : \text{O} = 18.1 : 16$; the equivalent of aluminium is 9.03; hence the formula of alumina might be:

$$\begin{array}{ccccccc} \text{AlO} & & \text{Al}_2\text{O}_3 & & \text{AlO}_2 & & \text{AlO}_3 \dots \\ \text{O : Al} \dots 16 : 18.1 & 48 : 27.1 & 32 : 36.2 & 48 : 54.2 \dots \end{array}$$

that is, the atomic weight of aluminium might be 18.1, 27.1, 36.2, 54.2 . . . There is nothing in the composition of the oxide to show which of these numbers should

be selected. It is known, however, that ferric oxide— Fe_2O_3 —forms a series of iron alums isomorphous with the aluminium alums; hence it is inferred that the constitution of aluminium oxide is Al_2O_3 , like that of ferric oxide, and that the atomic weight of aluminium is 27.1.

Questions.

1. State the law of isomorphism, and give examples of its utility in fixing the atomic weights of elements.—*St. Andrews Univ.*

2. "Sodium sulphite, $\text{Na}_2\text{SO}_3 \cdot 7\text{H}_2\text{O}$, forms colourless *monoclinic crystals*, with a *saline taste*, and *alkaline reaction*. The crystals *effloresce* in dry air; they are *soluble* in water; and become *anhydrous* when heated to 150° . At a higher temperature the crystals *decompose*, forming a yellow liquid." Explain the meaning of the terms in italics in this quotation.

3. "The element sulphur is *dimorphous*. It occurs in both *rhombic and monoclinic crystals*, and also in an *amorphous* condition." Explain the meaning of the terms in italics in this quotation.

4. Discuss the statement of Mitscherlich that equal numbers of atoms similarly combined exhibit the same crystalline form; identity of crystalline form being independent of the chemical nature of the atoms, but conditioned only by their number and configuration. Point out the precise value, as well as the limitations of this "law" as regards its bearing on chemical classification.—*Science and Art Dept.*

5. State the law of isomorphism; illustrate its application to the determination of the class relations of the elements, and point out apparent exceptions to the law.—*London Univ.*

6. Explain the meaning of the terms solution, saturated solution, and crystallization. Show how it can be proved experimentally that hot water is generally a better solvent than cold water.—*Cape Univ.*

7. If the solubility of sodium chloride in water at 20° is 35.6, how much water will be needed at this temperature for the solution of a kilogram of the salt?

CHAPTER XI

OZONE AND HYDROGEN PEROXIDE

§ 1. The Modes of Formation, and Preparation of Ozone.

OZONE. Molecular weight, $O_3 = 48$. Boiling point about -120° . Relative vapour density ($H_2 = 2$), 46.65; (air = 1) 1.62.

The formation of Ozone.—Ozone is produced by the action of the ultra-violet rays and radium radiations on oxygen. It is said to be formed by the violent mechanical disturbance of air as when grinding wheels are being tested for bursting speed; but the effect may here be due to the electrification of the air. Ozone can generally be detected in the oxygen gas obtained during the electrolysis of acidulated water (p. 33). By the electrolysis of sulphuric acid (between 1.075 and 1.1 specific gravity), with an anode made by imbedding platinum foil in glass and grinding away the edge so that a line of platinum, 0.1 mm. broad, is exposed, oxygen containing 17 to 23 per cent. of ozone has been obtained. The yield increases with increasing current density. The presence of ozone can be shown by means of a strip of paper wetted with a solution of starch and a little potassium iodide. The paper so prepared becomes blue with ozone, not with oxygen.

The oxygen liberated by many reactions also contains ozone. For instance, manganese dioxide and sulphuric acid; barium dioxide and sulphuric acid; potassium permanganate and sulphuric acid, persulphuric acid, persulphates, etc. Purified potassium chlorate gives oxygen free from ozone, but if traces of some indifferent substances be present, ozone may be formed. When fluorine decomposes water forming hydrogen fluoride and oxygen, from 13 to 14 per cent. of the "oxygen gas" is ozone.

Ozone is formed during the slow oxidation of many substances. Ozone can be detected in the atmosphere of a flask containing a couple of sticks of clean phosphorus. J. H. van't Hoff expressed the opinion that the effects attributed to ozone may possibly be due to neutral or charged atoms of oxygen. The ozonized air is obtained by aspirating the air through a flask containing clean pieces of phosphorus partially submerged in water. Ozone is also said to be formed when turpentine, several hydrocarbons, coal tar, and many essential oils are oxidized. J. K. Bocke thinks that the effect is due to the formation of nitrogen oxide. Ozone is said to be formed during the combustion of ether as well as during the combustion of hydrogen compounds generally. At any rate, the potassium iodide test indicates the formation of ozone (or hydrogen peroxide) when a spiral of hot platinum is placed above the surface of a little ether in the bottom of a beaker. The ether burns on the surface of the platinum, and the platinum remains

incandescent as long as any ether remains in the beaker. Again, if a strong jet of air be blown through the upper portion of a Bunsen's flame; or if a very narrow tube be presented to the lower edge of a Bunsen's flame and a slow current of air be aspirated through a little potassium iodide dissolved in water, the ozone (or hydrogen peroxide) reaction is obtained.

Ozone seems to be fairly stable at ordinary temperatures, although it gradually decomposes on standing. It also appears to be fairly stable at high temperatures, while at intermediate temperatures it is unstable. Hence ozone may be formed by heating oxygen to a high temperature, and, by suddenly chilling the gas, it can be cooled below the temperature at which it is very unstable without being all decomposed. The chilling can

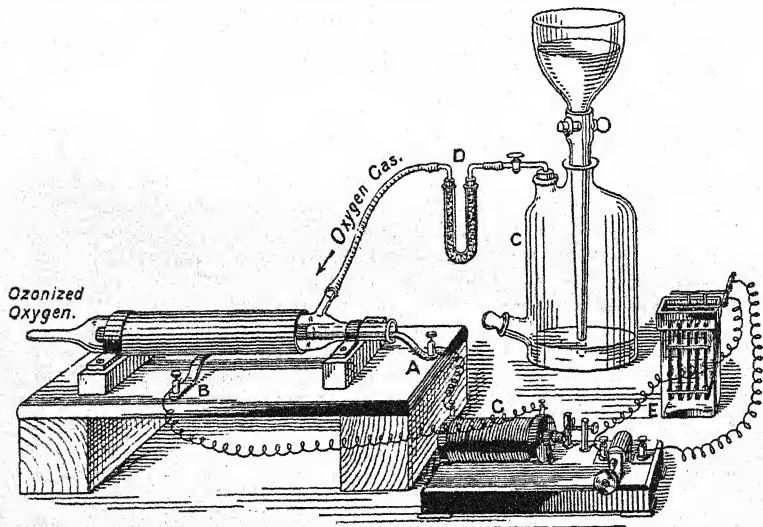


FIG. 80.—Preparation of Ozone with Siemens' Tube.

be done by blowing air or oxygen against the hot pencil of a Nernst's burner; or by dipping a hot Nernst's pencil, or hot platinum wire in liquid air.

Method of preparing ozone.—The most usual method of making ozone, or rather of preparing ozonized air and ozonized oxygen is to expose air or oxygen to a silent discharge of electricity. Quite a number of instruments are available. That illustrated at *AB*, Fig. 80, is virtually the one devised by W. von Siemens (1858). *Siemens' ozonizer* consists of two concentric tubes. The inner tube is coated on its inner surface with tinfoil in metallic contact with the terminal *A*; and the outer tube is coated on its outer surface with tinfoil in metallic contact with the terminal *B*. The two terminals are connected with an induction coil. A slow stream of oxygen is led from the gasholder *C* through the calcium chloride drying tube *D*, and then through the annular space between the concentric tubes, and is there exposed to the action of the silent discharge of electricity, operated by the induction coil *G*, and battery *E*. The gas issuing from the ozone

tube or ozonizer is charged with 3 to 8 per cent. of ozone. In *Brodie's* or *Berthelot's* ozone tube the tinfoil coatings are replaced with sulphuric acid (Fig. 84). If air be used in place of oxygen, some nitrogen oxides are said to be formed at the same time. The presence of moisture is said to reduce the yield of ozone,¹ although no difference has been detected in the amount of decomposition of the dry and moist gas when heated for some time at 100°.

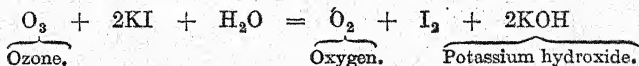
The formation of ozone by the evaporation of water or dew is said to account for the bleaching of linen spread on lawns.

Much atmospheric ozone seems to be formed by the action of solar ultra-violet light in the upper regions of the atmosphere. Ozone is here present in relatively greater proportions; thus, on the Alps at an altitude of 20 kilometres the air had 5 parts of ozone per million (volumes); and at an altitude of 2·1 kilometres, half this proportion. Electrical discharges in the atmosphere also produce ozone. Homer in several passages referred to the odour which attended the "thunderbolts" or flashes of lightning.

§ 2. Properties and Occurrence of Ozone.

Ozonized air has a strong unpleasant smell. The smell reminds some people of sulphur dioxide, others of garlic, and others of chlorine. If air highly charged with ozone be breathed for any length of time, it produces headache; but in minute quantities the odour is pleasing and refreshing. Ozone is slightly soluble in water—100 volumes of water at ordinary temperatures and pressure dissolve about one volume of ozone—and the water smells of the ozone and exhibits many of the properties of ozone. The water slowly reacts with the ozone.² Ozone is dissolved by essential oils—turpentine, thyme oil, cinnamon oil, etc. The unsaturated carbon compounds with a double bond between carbon atoms absorb ozone quantitatively, forming *ozonides*, whereas those with a triple-bonded carbon atom do not absorb ozone in the same way. Thus, butylene, C_4H_8 forms the two ozonides $C_4H_8O_3$ and $C_4H_8O_4$.

Action on potassium iodide.—Unlike oxygen, ozone liberates iodine from neutral potassium iodide. This can be shown by dipping paper in a solution of potassium iodide and holding it at the exit tube of the ozonizer. The paper turns brown owing to the liberation of iodine. If a little starch be mixed with the potassium iodide, the paper will appear blue. The reaction is usually represented:



The potassium hydroxide is alkaline. Hence if red litmus paper be steeped in water containing a trace of potassium iodide, the moist paper, when exposed to ozonized air, will be coloured blue owing to the action of the potassium hydroxide on the red litmus. If ozone acts upon an acidified solution of potassium iodide, the result is different from that which occurs with neutral potassium iodide. The action also depends upon the temperature and the strength of the solution (B. Brodie, 1872).

¹ Ozone is decomposed by cork and indiarubber. In consequence, these materials should not be used for any part of the ozonizer in contact with the gas.

² The so-called "ozone water" is largely a mixture of nitrogen oxide, chlorine, hypochlorous acid, and calcium hypochlorite.

The method used for the determination of ozone in air, etc., is based upon the reaction symbolized above. A known volume of air is drawn through a neutral solution of potassium iodide, and the liberated iodine is determined by acidifying the solution, and titrating the liberated iodine with a standard solution of sodium thiosulphate, as will be indicated later. The standard method for estimating ozone in, say, the atmosphere is to expose ozone test papers ¹ for a definite time to the air and compare the resulting tint with a standard scale of colours obtained with air containing known quantities of ozone. Since other oxidizing substances, likely to be present in air, produce a similar reaction, it is a moot question whether the large number of "ozone determinations" which have been made really represent ozone, or hydrogen peroxide, nitrogen oxides, chlorine, etc. Rather does the result of the test represent the presence of "oxidizing substances."

Oxidizing action of ozone.—Ozone is a very powerful oxidizing agent. Organic matter—cork, rubber, etc.—is rapidly corroded by ozone. Paper coloured by a solution of indigo sulphate or a solution of litmus is bleached. Other substances—hydrogen peroxide, chlorine, nitrogen oxide, etc.—give these reactions and hence the tests do not distinguish between ozone and the compounds just named. Put a globule of mercury in a small flask, pass ozonized air into the flask and shake the globule of mercury about. The mercury loses its lustre, and spreads a film over the walls of the flask. The globule of mercury is restored when the film is shaken up with water. If a piece of silver foil cleaned with silver sand be heated in a Bunsen's burner for a moment, and while still warm, held in a stream of ozonized air, the silver is blackened, owing, it is said, to the formation of silver peroxide. Silver perfectly free from the oxide does not show the reaction at ordinary temperatures unless the metal be contaminated with some substance like oxide of silver, cobalt, nickel, etc., which acts as a catalytic agent. A trace of oxide is supposed to be formed when the silver is heated in the Bunsen's burner. The reaction is characteristic of ozone, but it is not very sensitive. Ozone converts dark brown coloured lead sulphide into white coloured lead sulphate: $\text{PbS} + 4\text{O}_3 = \text{PbSO}_4 + 4\text{O}_2$. This can be shown by holding a strip of paper, which has been steeped in a solution of lead acetate and "browned" by hydrogen sulphide, in a stream of ozonized oxygen or ozonized air. Many other sulphides—copper, antimony, zinc, cadmium—behave in a similar manner. Nickel and cobalt sulphides form peroxides and sulphuric acid. A solution of manganese sulphate used as an invisible ink is browned by exposure to ozone.

The action of ozone on peroxides.—Ozone decomposes in the presence of finely divided platinum, lead dioxide, manganese dioxide, silver and copper oxides. The ozone is converted into ordinary oxygen without decomposing the oxides. Hence the reactions are grouped among catalytic reactions. The effect can be shown by passing ozonized air through a tube containing copper oxide and testing the issuing gas by ozone test paper. No indication of ozone is obtained. When ozone is brought into contact with sodium peroxide, the two substances mutually decompose and oxygen is liberated: $\text{O}_3 + \text{Na}_2\text{O}_2 + \text{H}_2\text{O} = 2\text{NaOH} + 2\text{O}_2$.

¹ Papers steeped in an emulsion of starch containing a small proportion of potassium iodide are called ozone test papers.

The action of the silent discharge on ozone.—The silent discharge has a deozoneing as well as an ozoneing effect on oxygen. The speed of the ozoneization is proportional to the amount of oxygen present, and the speed of the deozoneization is proportional to the amount of ozone present. In other words, the reaction follows the law of opposing reactions. If the discharge be passed for an indefinite time, a certain definite limiting concentration of ozone will be reached when the rate of decomposition is equal to the rate of formation of the ozone: $3O_2 \rightleftharpoons 2O_3$. The greater the pressure of the gas, the greater the yield; and if the density of the gas be kept constant, the yield of ozone is not very different at temperatures between 10° and 80° . The presence of moisture also reduces the yield. (E. Warburg, 1906.)

The action of heat on ozone.—Ozone decomposes slowly at ordinary temperatures, and the rate of decomposition is increased by raising the temperature. According to E. Warburg (1902), at 16° , one per cent. of pure ozone per litre of oxygen decomposes in 1.7 minutes; at 100° , in 0.003 minute; at 127° , in 0.0027 minute; and at 1000° the decomposition is almost instantaneous, for 0.0007 second suffices for the decomposition of $\frac{999}{1000}$ of the ozone present. Hence to show the formation of ozone at high temperatures, the velocity of cooling must be greater than the rate of decomposition. If a tube be attached to the exit tube of the ozonizer so that a slow current of ozonized oxygen can be passed through the tube heated to about 300° , the issuing gas will give no reaction for ozone with the ozone test papers. Since ozone is formed at very high temperatures, there must be a reversal in the stability, before the temperature of formation is reached.

The effect of cooling ozone.—By passing ozonized oxygen through a tube cooled by immersion in boiling liquid oxygen, or by ozonizing oxygen in a tube kept at this temperature, a solution of ozone in liquid oxygen is obtained. By allowing the liquid to boil, most of the oxygen is removed. In this manner a deep indigo blue liquid is obtained which is opaque in layers 2 mm. thick. The liquid is said to be explosive. By allowing the blue liquid to vaporize, A. Ladenburg (1898) obtained a gas containing about 86 per cent. of ozone; the liquid is particularly liable to explode when it reaches -120° the boiling point of ozone.

The effusion of gases.—T. Graham (1832) found that the law of diffusion, indicated on p. 127, holds good for the passage of a gas through a very fine aperture in a metal plate. Graham called the phenomenon the effusion of gases. The speed of effusion is therefore inversely as the square root of the specific gravity. If a gas of density D_1 flows out of the tube in the time t_1 , and another gas of density D_2 flows out in the time t_2 , then, according to the law of effusion: $D_1 : D_2 = t_2^2 : t_1^2$. R. Bunsen (1857)

utilized this fact to determine the specific gravity of a gas when but a small quantity of the gas is available. In N. H. Schilling's effusion

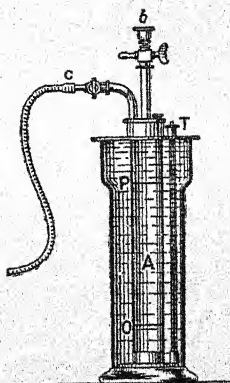


FIG. 81.—Schilling's Effusion Apparatus.

apparatus (1879) the gas is introduced into a glass tube, luted to a brass cover, *via* the cocks *b*, *c*, Fig. 81. This tube is placed in a cylinder filled with water and the temperature is indicated by the thermometer *T*. The perforated platinum plate is fixed at *b*. The time taken for the water to press the gas from the level *O* to the level *P* on the cylinder is noted. The experiment is repeated with another gas of known density. A. Ladenberg (1898) used this apparatus to determine the specific gravity of ozone.

EXAMPLE.—A. Ladenberg (1898) found that a mixture containing 86.16 per cent. of ozone required 367.5 seconds to effuse under conditions where pure oxygen required 430 seconds. Hence, determine the specific gravity of ozone. From the example on p. 128, it follows that the specific gravity of the ozonized oxygen is 1.3698, oxygen = 1. From the example on p. 32, it follows that the specific gravity of ozone is 1.456, if oxygen be unity; and 46.6 if oxygen O_2 be 32. Note that this does not establish the molecular weight by Avogadro's hypothesis, because, in determining the proportion of ozone in the mixture, by estimating the amount of iodine liberated by a given volume of the gas, it was assumed that the reaction proceeds as indicated in a preceding equation which in turn assumes that the formula of ozone is O_3 .

Occurrence of ozone.—Some samples of fluorspar are said to contain traces of ozone. Ozone is supposed to occur in small quantities in the atmosphere near the seaside where it is supposed to be formed by the evaporation of water. According to A. Houzeau, country air contains about one volume of ozone per 700,000 volumes of air. The maximum amount of ozone in the atmosphere is said to occur during the spring months, and gradually diminishes, reaching a minimum in winter. Ozone is absent in the air of towns, dwelling-houses, over marshes, and wherever organic matter is present. It has been shown that much of the alleged ozone may not be ozone at all. This is due to the imperfection of the tests employed (p. 221). T. Andrews found that "oxidizing matter" in the air was destroyed by heating the air to 260° . This would not be the case if the oxidizing matter was chlorine or nitrogen oxide. But we shall see very shortly that Andrews' test does not discriminate between ozone and hydrogen peroxide. Hence there is some uncertainty about many of the reactions where ozone is supposed to be formed; and also in many reports of the proportion of ozone in the atmosphere.

The distinction between ozone and hydrogen peroxide.—C. Arnold and C. Mentzel (1902) propose the use of test papers soaked in an alcoholic solution of "tetramethyl base." These are not affected by hydrogen peroxide, but are turned violet with ozone; blue, with chlorine and bromine; and straw yellow, with nitric oxide. C. Engler and W. Wild (1896) state that if a mixture of ozone and hydrogen peroxide be passed through a concentrated solution of chromic acid, the hydrogen peroxide is alone decomposed, the ozone is not affected. The same investigators say that paper steeped in a concentrated aqueous solution of manganous chloride is turned brown by ozone, but not by hydrogen peroxide.

Uses.—Ozone is used for the purification of water in special cases, but is rather expensive. Its function is to oxidize the organic matter, and sterilize the water. Ozonized air is also used in ventilation for bleaching oxidizing oil in the manufacture of linoleum, etc. The industrial applications have stimulated inventors, and accordingly, a number of fairly efficient ozonizers have been placed on the market. In most of these, a

high tension alternating electric discharge is sent across a space through which the air to be ozonized passes.

§ 3. The History and Constitution of Ozone.

1. The discovery of ozone by Schönbein (1840).—In 1775 M. van Marum said that he noticed a peculiar smell in the vicinity of electrical machines in motion, but he does not appear to have made any attempt to find the cause of the smell. W. Cruickshank also, in 1801 mentioned a similar smell attending the electrolysis of water. C. F. Schönbein (1839–80) was the first to recognize that the smell was due to the formation of a substance to which he gave the name ozone—from the Greek $\omega\zeta\omega$ (ozo), I smell. According to Schönbein, ozone is a distinct form of matter with an identity of its own. Schönbein said that the same substance was produced when an electrical machine is working; when water is electrolyzed; and when moist air is passed over oxidizing phosphorus. The nature of ozone was the subject of much discussion in Schönbein's day. Schönbein first considered it to be a new elementary body which was a component of the nitrogen in the atmosphere.

2. Is ozone a condensed form of oxygen or an oxide of hydrogen?

—C. Marignac, A. de la Rive, and others (about 1845) showed that moist silver, when exposed to ozone, forms silver oxide, and that potassium iodide—KI—can be oxidized to potassium iodate— KIO_3 . This narrowed the question, for it appeared that ozone is either (1) a form of matter identical with oxygen (C. Marignac, A. de la Rive); or (2) oxidized

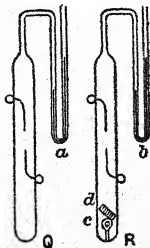


FIG. 82.—Andrews and Tait's Ozone Tubes.

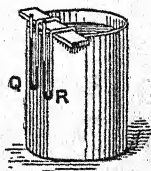


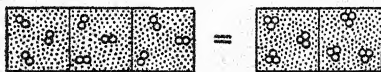
FIG. 83.—Andrews and Tait's Experiment.

water, that is, a peroxide of hydrogen (Schönbein). The hydrogen oxide theory was not given up until 1860, when T. Andrews and P. C. Tait proved that if an electric discharge—silent or spark—be passed through pure dry oxygen, a contraction occurs amounting to $\frac{1}{12}$ of the original volume. The oxygen was sealed in a tube, shaped as indicated in Fig. 82, Q, and subjected to the silent discharge, *vid* the platinum wires sealed into the glass. In Andrews and Tait's experiment, the contraction in volume was measured by attaching to the tubes a small manometer, *a* and *b*, charged with concentrated sulphuric acid. A duplicate tube *Q*, Fig. 83, containing air was treated along with the tube containing the oxygen, *R*, Fig. 82, so that any changes due to variations of pressure or temperature during the experiment could be corrected. The tubes during the experiment, were placed in a water-tank, as indicated in Fig. 83, to keep the temperature uniform. When ozonized oxygen is heated to 270° , and allowed to cool, the original volume of oxygen is obtained; and when a thin glass bulb, *c*, Fig. 82, *R*, of potassium iodide is sealed in the tube along with the oxygen, and after ozonization, broken by shaking the bulb against a piece of glass tubing *d*, iodine is liberated without any perceptible change in volume. If the gas

which has been treated with potassium iodide be heated to 270° as before, no change in volume can be detected. Hence Andrews and Tait concluded that ozone is a condensed form of oxygen. This, however, tells us nothing about the weight of oxygen in a given volume of ozone, *i.e.* the number of atoms in the molecule of ozone.

The absence of hydrogen in ozone was inferred by Soret (1863) from an experiment in which ozone was thoroughly dried, and then decomposed by heat. No trace of any compound of hydrogen—*e.g.* water—could be detected in the products of decomposition. Hence it is inferred that ozone is not a compound of hydrogen with oxygen; ozone contains nothing but oxygen. A similar experiment was made by Schönbein in 1849, but its importance does not appear to have been appreciated at that time.

3. Ozone is a form of oxygen in which three volumes of oxygen are condensed to two volumes.—Since the volume of ozonized oxygen undergoes no change when mixed with a solution of potassium iodide, it is inferred that the oxidation of potassium iodide can only be effected by so much oxygen in ozone as has been condensed with ordinary oxygen to form ozone. This excess of oxygen is absorbed by the solution of potassium iodide, and the ordinary oxygen which remains has the same volume as the ozone present before the action of the potassium iodide. Hence no new contraction occurs with potassium iodide. In symbols, the formula for ozone is O_{2+n} . Andrews and Tait did not determine the numerical value of n . The formula for ozone might be O_3 , O_4 , O_5 , . . . The special difficulty involved in this determination arises from the fact that ozone cannot be obtained free from oxygen; and, accordingly, the regular methods of determining the molecular weights—vapour density, etc.—cannot be applied. W. Odling (1861) proposed to take the simplest possible formula O_3 , thus assuming that three volumes of oxygen are condensed to form two volumes of ozone. Interpreting this unproved assumption in the light of Avogadro's hypothesis, p. 75, we have:



3 Volumes oxygen give 2 Volumes ozone.

This beautiful hypothesis, said C. W. Heaton, although accounting perfectly for all the known facts, was yet but a probability. One link was lacking in the chain of evidence, and that link Soret supplied by a happily devised experiment. In 1866 he took advantage of the fact, known to Schönbein, that essential oils absorb ozone without decomposition. Hence, if ozonized oxygen be shaken with, say, cinnamon oil, the ozone will be removed from the mixture. On treating one portion of a sample of ozonized oxygen with cinnamon oil, Soret found that a contraction corresponding with about 2 c.c. was obtained. On heating another portion of the same sample so as to convert the ozone into ordinary oxygen, Soret found an expansion corresponding with 1 c.c. Hence it was inferred that three volumes of oxygen produce two volumes of ozone. Soret's work was rather crude, but B. Brodie (1872) repeated the experiments with cinnamon oil, turpentine, stannous chloride, in such a way that the above conclusion was the only possible interpretation of the experiments.

Many neat ways of illustrating the volume relations of oxygen and ozone have been devised. G. S. Newth's apparatus (1896), slightly modified, consists of two concentric tubes, Fig. 84. The inner tube has a hollow stopper ground to fit the outer tube. It contains dilute sulphuric acid. The inner tube has two little projections *A*, and the outer tube has three projections, *B*, in such a position that a sealed thin glass tube containing cinnamon oil can be broken, when desired, by twisting the stopper or the inner tube. The outer tube is fitted with a three-way cock *D*, connected with a manometer charged with concentrated sulphuric acid. The apparatus is placed in a cylinder containing, say, ice and water. The annular space between the two tubes is filled with oxygen, *via* the cocks *E* and *D*. The manometer is then put in communication with the annular space between the two tubes. Note the level of liquid in the manometer. Pass a current from an induction coil, so as to ozonize the oxygen sufficiently to give, say, a one centimetre contraction on the manometer. Note the contraction. Give the stopper a twist so as to break the glass tube containing the cinnamon oil, the contraction which occurs will be twice the former contraction, namely 2 cm. more.¹ The same apparatus can be employed for showing that no contraction occurs when ozone is treated with potassium iodide by using a tube *C* with this substance in place of cinnamon oil.

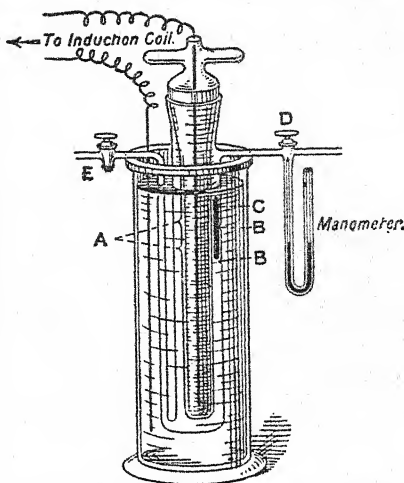


FIG. 84.—Newth's Apparatus (modified).

The formula for ozone— O_3 —obtained by many modifications of this experiment has been confirmed by the application of Graham's diffusion test, p. 223. Ozone is 1.5 times as heavy as an equal volume of oxygen.

4. The graphic formula for ozone.—Ozone is generally symbolized graphically: $\begin{smallmatrix} O \\ \diagup \quad \diagdown \\ O-O \end{smallmatrix}$, $\begin{smallmatrix} O \\ \diagup \quad \diagdown \\ O \equiv O \end{smallmatrix}$, or $O \equiv O \Rightarrow O$. Owing to the peculiar oxidizing

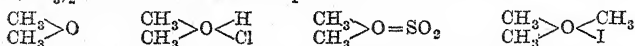
qualities of the odd oxygen atom in the molecule some consider that the three oxygen atoms are not symmetrically placed in the molecule, whereas in the above formula they apparently all have equal value. Ozone is often considered to be a compound of one quadrivalent oxygen atom with two bivalent oxygen atoms, and the formula is accordingly written: $O=O=O$; or, if two quadrivalent oxygen atoms are united with one

bivalent oxygen atom: $\begin{smallmatrix} O \\ \diagup \quad \diagdown \\ O \equiv O \end{smallmatrix}$; hence, the graphic formula for ozone is still *sub judice*.

Belief in the existence of quadrivalent oxygen atoms is mainly based on evidence deduced from organic chemistry. For instance, hydrogen chloride,

¹ It may be advisable to level the liquid in the manometer, after the first contraction, before breaking the capillary tube, by admitting either air or oxygen.

sulphur dioxide methyl iodide, etc., can be made to combine directly with methyl ether $(\text{CH}_3)_2\text{O}$ to form a series of compounds :



and also a series of oxonium salts. These compounds are not discussed in inorganic chemistry. Evidence in favour of a sexivalent oxygen atom has been brought forward; the so-called ozonic acid, H_2O_4 , of A. Bach (1902), and other organic compounds might be cited in illustration. Hence oxygen, though usually bivalent, may be quadrivalent and probably also sexivalent. There is some evidence indicating the formation of an even more condensed form of oxygen than ozone when the latter is produced by a silent electrical discharge of very high voltage—18,000 volts. The last fractions of the fractional distillation of the liquefied product furnished Harries with a hyperozone, but it has not been isolated.

§ 4. Hydrogen Peroxide or Hydrogen Dioxide—Formation and Preparation.

Molecular weight, $\text{H}_2\text{O}_2 = 34.02$. Melting point, -2° ; boiling point, $84^\circ\text{--}85^\circ$ at 68 mm. pressure; and at 26 mm. pressure, $68^\circ\text{--}69^\circ$. Specific gravity, 1.4996.

Hydrogen peroxide or *oxygenated water* is a remarkable compound which has attracted the attention of chemists from the time of its discovery by L. J. Thénard in 1818. Hydrogen peroxide is formed when oxygen is bubbled about the electrode from which hydrogen is being evolved during the electrolysis of dilute acid, p. 33. The formation of hydrogen peroxide at the anode during the electrolysis of dilute sulphuric acid by a current of high density is considered by M. Traube to be the result of a secondary reaction, the direct reaction is supposed to furnish persulphuric acid, $\text{H}_2\text{S}_2\text{O}_8$. Water confined in a quartz vessel is decomposed by exposure to ultraviolet light rays—from a mercury lamp, sunlight, etc.—and hydrogen peroxide and hydrogen are formed: $2\text{H}_2\text{O} = \text{H}_2\text{O}_2 + \text{H}_2$. Hydrogen peroxide is produced during the combustion of hydrogen in air. For instance, when a jet of burning hydrogen impinges on the surface of cold water in which ice is floating, or on ice itself, hydrogen peroxide can be detected in the water. Hydrogen peroxide is formed when moist ether is exposed to sunlight; and when ozonized oxygen or air is passed through water on the surface of which a little ether floats. If a little water is placed in a beaker containing ether, and the latter is burnt by placing a spiral of hot platinum wire just over the surface of the liquid (p. 221), hydrogen peroxide can be detected in the water after all the ether has burnt away. It is supposed that the ozone first produced forms a peroxide with the ether, and that this is decomposed by the water reforming ether and hydrogen peroxide.¹

Like ozone, hydrogen peroxide can be formed at a high temperature by passing a current of moist oxygen through a tube at about 2000° and rapidly chilling the issuing gases. In H. St. C. Deville's "hot and cold tube" method of conducting the experiment, a narrow silver or platinum tube is kept cool by a current of cold water. This tube is placed in the centre of a porcelain tube, Fig. 85. A current of gas is passed along the annular space between the two tubes. This arrangement is placed in a furnace so that the gas can be heated to a very high temperature. The products of decomposition are suddenly chilled by the

¹ The ether is thus regarded as a catalytic agent, and the reaction is supposed to proceed by a set of consecutive reactions, p. 163.

cold tube and partially prevented from recombining as they are carried out of the hot zone. The products of many high temperature reactions can thus be examined at ordinary temperatures.

Hydrogen peroxide is often formed when a substance is oxidized in the presence of moisture. For instance, when zinc, copper, or lead is shaken up with air and dilute sulphuric acid (1 : 55), the reaction symbolized : $\text{Zn} + 2\text{H}_2\text{O} + \text{O}_2 = \text{Zn}(\text{OH})_2 + \text{H}_2\text{O}_2$; and $\text{Zn}(\text{OH})_2 + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + 2\text{H}_2\text{O}$ occurs. It will be observed that twice as much oxygen is required for the oxidation process as is actually consumed in oxidizing the zinc: $\text{Zn} + \text{O}_2 = \text{ZnO} + \text{O}$; $\text{H}_2\text{O} + \text{O} = \text{H}_2\text{O}_2$. One half of the oxygen is said to be used in the primary process and the other half in the secondary reactions. The reaction belongs to the type of concurrent or side reactions discussed on p. 166, but since half a molecule of oxygen is used in each, the two concurrent reactions are not independent of one another. This particular type of reaction is known as auto-oxidation. There is a considerable difference of opinion as to the mechanism of auto-oxidation. The oxygen used in the secondary reaction—formation of hydrogen peroxide—is said to be “rendered active” by the primary reaction. The formation of ozone during the oxidation of phosphorus (p. 221) is another example.

By treating a cold aqueous solution of sodium peroxide with dilute and cold hydrochloric acid, a solution of hydrogen peroxide mixed with sodium chloride is obtained: $\text{Na}_2\text{O}_2 + 2\text{HCl} = 2\text{NaCl} + \text{H}_2\text{O}_2$; and by treating potassium peroxide, K_2O_4 , with tartaric acid in the cold, an aqueous solution of hydrogen peroxide contaminated with a little potassium tartrate is obtained, and oxygen is evolved. Most of the potassium tartrate separates from the cold solution.

Barium peroxide (p. 159) is the usual starting-point for the preparation of hydrogen peroxide. Gradually add barium peroxide to ice-cold water through which a stream of carbon dioxide is passing. The insoluble barium carbonate is precipitated, and a dilute aqueous solution of hydrogen peroxide remains: $\text{BaO}_2 + \text{CO}_2 + \text{H}_2\text{O} = \text{BaCO}_3 + \text{H}_2\text{O}_2$. If an excess of carbon dioxide be used, the yield of hydrogen peroxide is low and an insoluble barium percarbonate, BaCO_4 , is precipitated. Or barium peroxide, mixed with a little ice-cold water is gradually added to ice-cold dilute hydrochloric, sulphuric, silico-fluoric, or phosphoric acid. A barium salt—chloride, sulphate, silicofluoride or phosphate—and hydrogen peroxide are formed. In the first case, the barium chloride is soluble. It can be removed by adding just sufficient silver sulphate to precipitate insoluble barium sulphate and silver chloride: $\text{BaCl}_2 + \text{Ag}_2\text{SO}_4 = \text{BaSO}_4 + 2\text{AgCl}$. This method is of historical interest because J. Thénard employed a similar process when he discovered hydrogen peroxide in 1818. The sulphuric acid process is more commonly used. In this, insoluble barium sulphate is precipitated. If concentrated sulphuric acid

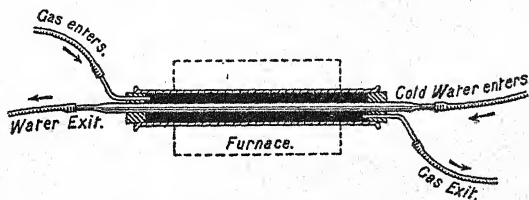


FIG. 85.—Deville's "Hot and Cold Tube."

be allowed to react with barium dioxide, ozonized oxygen is evolved; a more dilute acid gives oxygen; and a very dilute acid gives hydrogen peroxide. The technical preparation of hydrogen peroxide from potassium persulphate or persulphuric acid is indicated later, p. 545.

The preparation of almost pure hydrogen peroxide.—Barium peroxide, suspended in a little water, is gradually added to a mixture of equal volumes of water and sulphuric acid (cooled by a freezing mixture of ice and salt) until the solution is just barely acid. If too much barium peroxide has been added, a little more sulphuric acid is needed. Keep the solution in a freezing mixture for about a day. Filter off the insoluble matter, and evaporate the liquid on a water-bath, at about 70° , in a smooth platinum or porcelain basin until signs of effervescence appear. This will occur when the solution contains about 45 per cent. of hydrogen peroxide. Cool the solution quickly. Concentrated solutions soon decompose if they are not kept cold.

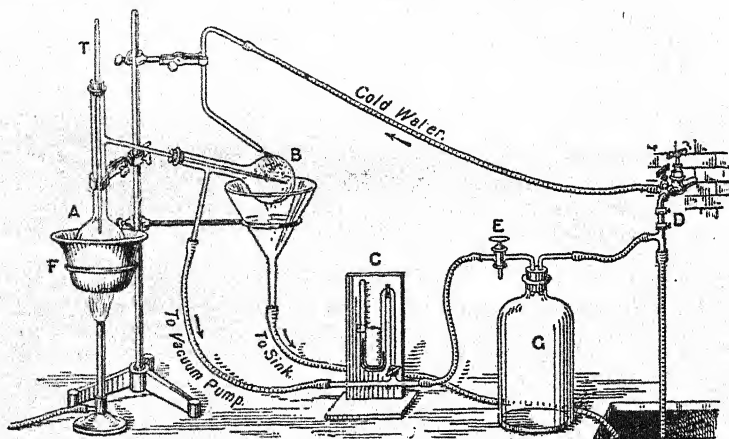


FIG. 86.—Distillation of Hydrogen Peroxide under Reduced Pressure.

A solution of hydrogen peroxide decomposes rapidly when heated to 100° —even if the solution be dilute—hence, for a long time, the concentration of an aqueous solution of hydrogen peroxide by fractional distillation was thought to be impossible; but R. Wolfenstein (1902) discovered that the hydrogen peroxide can be distilled under reduced pressure without undue decomposition, p. 188.

To distil the solution under reduced pressure, fit up the apparatus indicated in Fig. 86. Transfer the solution to a round-bottom thick-walled litre flask A, fitted with a receiver B and a thermometer T, by means of one-hole rubber stoppers. The receiver B is placed over a funnel so that cold water can be sprayed on the receiver, and run off to the sink through rubber tubing attached to the stem of the funnel. The side neck of the receiver is connected, by pressure tubing, with a manometer; which in turn is connected with a 3-way stopcock E, a water trap G, and a filter pump D. The flask A is heated by an oil bath F, and Bunsen's burner. When the manometer shows a pressure of about 15 mm. and the thermometer a temperature between 35° and 40° , a dilute aqueous solution of hydrogen

peroxide in water distills into the receiver. The temperature rises gradually to about 70° when a very concentrated solution of hydrogen peroxide remains in the distilling flask A. Further concentration is best effected by placing a beaker containing some of the hydrogen peroxide solution in a mixture of solid carbon dioxide and ether. The whole mass freezes. Drop a little of the frozen solid into a portion of the concentrated hydrogen peroxide solution. At between -8° to -10° small needle-shaped crystals separate. Drain away the mother liquor from the crystals; melt the crystals and cool the mass, so that another crop of crystals is obtained. Repeat the operations.

In this way, it is possible to prepare 100 per cent. hydrogen peroxide. The solution remaining in the distilling flask will serve for most experiments where concentrated solutions of hydrogen peroxide are required. This solution can also be further concentrated by evaporation over concentrated sulphuric acid *in vacuo*. This operation is conducted as follows: The dish containing the mixture rests on the perforated shelf of a desiccator, Fig. 87. The desiccator has a layer of concentrated sulphuric acid below the perforated shelf. The lid of the desiccator, well greased, is placed in position. The desiccator is then exhausted by connecting the stoppered tube with the air pump. Any water vapour given off by the solution in the dish is gradually absorbed by the concentrated acid.

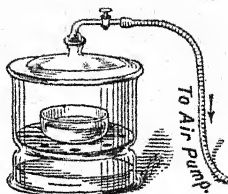


FIG. 87.—Evaporation under Reduced Pressure.

Concentrated hydrogen peroxide begins to attack the glass distilling flask at about 80° . Wolfenstein, by changing the receiver, obtained a solution of 90 per cent. hydrogen peroxide at 81° – 85° at 68 mm. pressure; and by repeatedly redistilling the product he got a liquid containing 99 per cent. of hydrogen peroxide and boiling at 84° – 85° C. at 68 mm. pressure.

§ 5. Hydrogen Peroxide—Occurrence and Properties.

Occurrence.—There is a similar uncertainty about the alleged occurrence of hydrogen peroxide in rain, snow, dew, and air as was indicated in dealing with the occurrence of ozone in air. Much of the published work does not clearly discriminate hydrogen peroxide from other oxidizing substances.

Properties.—Pure hydrogen peroxide is a viscid liquid; colourless, when viewed in thin layers, but with a bluish tinge when viewed in thick layers. The liquid has no smell. Dilute aqueous solutions have a peculiar bitter metallic taste. When a drop of liquid peroxide comes in contact with the skin, it forms a white blister. If concentrated sulphuric acid be mixed with hydrogen peroxide at a low enough temperature to prevent heating, oxygen rich in ozone is evolved. The liquid decomposes rapidly when heated at ordinary atmospheric pressures, but under reduced pressure (p. 188) it can be readily distilled. It boils between 68° and 69° under a pressure of about 26 mm., and at 84° – 85° under a pressure of 68 mm. The liquid crystallizes in needle-like prisms at -2° . The liquid is soluble in water in all proportions.

Pure hydrogen peroxide is fairly stable. Dilute aqueous solutions keep fairly well—particularly if acid—a 3 per cent. solution suffered no appreciable change when kept a year. Alkaline solutions do not keep very

well. Impurities like silica, iron, manganese, and alumina lead to a more rapid decomposition. If alcohol or ether be present, the aqueous solutions are more stable. The strength of aqueous solutions is represented commercially by the number of volumes of oxygen which 1 c.c. of the solution will furnish on decomposition. Thus 1 c.c. of a "10-volume solution" will give 10 volumes of oxygen when decomposed. A 3 per cent. solution of hydrogen peroxide is very nearly a "10-volume solution"; a 6 per cent. hydrogen peroxide solution is nearly of "20-volume" strength, and so on. There is a possible ambiguity because if the solution were decomposed by potassium permanganate, $2\text{KMnO}_4 + 5\text{H}_2\text{O}_2 + 3\text{H}_2\text{SO}_4 = 5\text{O}_2 + 2\text{MnSO}_4 + \text{K}_2\text{SO}_4 + 8\text{H}_2\text{O}$, half the oxygen comes from the permanganate. So that with a 10-volume solution of hydrogen peroxide 5 volumes would come from the peroxide and 5 volumes from the permanganate. Such a reaction would be excluded. The most concentrated solution on the market is called *perhydrol*, and it contains about 30 per cent. of hydrogen peroxide, corresponding with a "100-volume" strength.

Decomposition by contact action.—Pure hydrogen peroxide is decomposed very rapidly if any dust be present. A little platinum black dropped into the solution causes an explosion. Similar remarks apply to finely divided gold, silver, and similar metals, as well as to powdered manganese dioxide. The action appears to be catalytic since the manganese dioxide, etc., remain at the end of the action unchanged in composition. A small trace of some extremely finely divided metals—colloidal platinum, gold, etc.—can accelerate the decomposition of large amounts of hydrogen peroxide. The action has been compared with that of yeast on a solution of sugar, and these "colloidal" metals have been given the name *inorganic ferments*. Since the action is more or less hindered or retarded by traces of hydrocyanic acid, mercuric chloride, etc., the inorganic ferment is said to be poisoned by these agents. Charcoal or magnesium mixed with a trace of manganese dioxide ignites immediately. Finely powdered iron or lead remain quiescent, but if a trace of manganese dioxide be present, the iron burns. A few drops of liquid hydrogen peroxide on a piece of cotton wool will make the cotton inflame. Similar results are obtained with aqueous solutions of hydrogen peroxide, but the action is much less vigorous. Rough surfaces have a disturbing effect on the stability of hydrogen peroxide—a concentrated solution is decomposed when poured on to a ground-glass surface. The presence of small quantities of some substance—e.g. manganese silicate—make the solutions more stable; these agents have been called *anticatalysts*, *negative catalysts*, or *retarders*. Acetanilide is particularly useful in preserving hydrogen peroxide solutions.

Oxidizing properties.—Hydrogen peroxide resembles ozone in its strong oxidizing qualities. It liberates iodine from solutions of potassium iodide (see p. 222). The reaction is accelerated by acetic and mineral acids, and particularly by ferrous sulphate. According to Schönbein, one part of hydrogen peroxide in the presence of 25,000,000 parts of water can be detected by a mixture of potassium iodide and ferrous sulphate. It converts lead sulphide into lead sulphate, as was also the case with ozone. Hence the use of hydrogen peroxide for cleaning oil paintings which have been darkened by the action of hydrogen sulphide

—sometimes present in the air of towns—upon the lead compounds in the paint. The brownish-black coloured lead sulphide is transformed into white coloured lead sulphate. Sometimes the treatment is satisfactory, and sometimes it spoils the picture. Hydrogen peroxide bleaches many organic colouring agents, *e.g.* litmus and indigo solutions, but it does not affect mercury and silver like ozone.

Dilute solutions of hydrogen peroxide are accordingly used for bleaching silk, feathers, straw, hair, ivory, teeth, etc., where more violent bleaching agents would injure the material. Instead of hydrogen peroxide an acidified solution of sodium peroxide is sometimes employed. The actions are similar. Since the products of the decomposition of hydrogen peroxide—water and oxygen—are harmless, it is also used medicinally as an antiseptic. Hydrogen peroxide is also used in analytical work for the oxidation of sulphites to sulphates; arsenites to arsenates; chromic salts to chromates; ferrous to ferric salts; nitrites to nitrates, etc. H. B. Baker and L. H. Parker showed that if hydrogen peroxide be removed from synthetic water, the water attacks sodium amalgam much more slowly than water containing, say, 1 part of the peroxide in 100,000.

Peroxidizing properties.—Hydrogen peroxide forms peroxides of the alkalis and alkaline earths when treated with the corresponding hydroxides. For instance, with barium hydroxide: $\text{Ba(OH)}_2 + \text{H}_2\text{O}_2 = \text{BaO}_2 + 2\text{H}_2\text{O}$. W. Spring (1895) pointed out that hydrogen peroxide behaves in these reactions like an acid (p. 167). In confirmation, if sodium carbonate be added to hydrogen peroxide the corresponding alkaline peroxide is formed and carbon dioxide is evolved: $\text{H}_2\text{O}_2 + \text{Na}_2\text{CO}_3 = \text{Na}_2\text{O}_2 + \text{CO}_2 + \text{H}_2\text{O}$; on the contrary, if the hydrogen peroxide be added to the solution of the carbonate, oxygen is evolved: $2\text{H}_2\text{O}_2 + \text{Na}_2\text{CO}_3 = \text{Na}_2\text{CO}_3 + 2\text{H}_2\text{O} + \text{O}_2$. The sodium carbonate acts as a catalytic agent in the latter case. It is not at all uncommon to find reactions progressing differently according to the way the substances are mixed together.

With titanium salts, hydrogen peroxide gives an orange-yellow coloration supposed to be due to the formation of pertitanic acid: $\text{TiO}_2 + \text{H}_2\text{O}_2 = \text{H}_2\text{O} + \text{TiO}_3$. The particular tint depends upon the amount of titanium present, and hence the reaction is used for the determination of the amount of titanium in various materials. The tint of a solution containing an unknown amount of titanium is compared with that of a similar solution containing a known quantity of titanium; and the amount in the unknown solution determined by simple rule of three. The reaction is also used as a test for hydrogen peroxide. It is said that one part of titanium in 1800 parts of water gives a deep yellow coloration, and one part in 180,000 a light yellow coloration. Cerium and vanadium salts give brick-red coloration, and uranium a bluish coloration, but these reactions are not so sensitive as with titanium salts.

With chromic acid, H_2CrO_4 , hydrogen peroxide forms a blue solution which begins to decompose immediately with the evolution of oxygen. Some consider the oxide formed to be a compound of perchromic acid, HCrO_4 , with hydrogen peroxide, say, HCrO_5 (*q.v.*). This is pure hypothesis; no such compound has been isolated. The blue-coloured peroxide, whatever it be, is much more soluble in ether than in water, so that if a mixture of chromic acid and hydrogen peroxide in a test-tube be shaken

with ether, a blue ethereal solution of the peroxide will float on the surface of the aqueous layer. The compound decomposes when the ether is evaporated. This reaction is used for the detection of chromates: "Add dilute sulphuric acid to the chromate solution, and shake up with ether and hydrogen peroxide. The separation of a blue ethereal layer indicates chromic acid." The necessary modification of the process for the detection of hydrogen peroxide will be obvious. It is said that this method will indicate one part of hydrogen peroxide in 80,000 parts of water.

Reducing properties.—Hydrogen peroxide appears to act as a reducing as well as an oxidizing agent. With ozone it forms oxygen and water: $O_3 + H_2O_2 = H_2O + 2O_2$; with silver oxide, metallic silver and oxygen: $Ag_2O + H_2O_2 = 2Ag + H_2O + O_2$. Similarly with gold oxide: $Au_2O_3 + 3H_2O_2 = 2Au + 3H_2O + 3O_2$; and lead dioxide, $PbO_2 + H_2O_2 = PbO + H_2O + O_2$. Curiously enough, in these reactions, the reducing agent is itself reduced—usually the reducing agent is oxidized during the reduction. The lead dioxide obtained when red lead is digested with dilute nitric acid dissolves very slowly, but if a few drops of hydrogen peroxide be added, all the lead dioxide dissolves in a few moments. The lead dioxide is reduced to lead monoxide by the hydrogen peroxide, and the product dissolves immediately in the dilute acid. This method is generally employed to hasten the solution of red lead in dilute acid prior to analysis. Hydrogen peroxide in alkaline solutions oxidizes manganese oxide, MnO , to manganese dioxide, MnO_2 ; but in acid solutions it reduces manganese dioxide to manganous oxide: $MnO_2 + H_2SO_4 + H_2O_2 = MnSO_4 + 2H_2O + O_2$. The reducing action of the hydrogen peroxide is only apparent. According to B. Brodie (1872) the oxides of silver, manganese, etc., have an atom of oxygen which is readily disengaged from its combination. Similarly, hydrogen peroxide readily parts with its odd atom of oxygen. Consequently, the atom of oxygen in hydrogen peroxide is supposed to *oxidize* the odd oxygen atom in the metallic peroxide.

A solution of potassium permanganate— $KMnO_4$ —acidified with sulphuric acid is rapidly reduced by hydrogen peroxide: $5H_2O_2 + 2KMnO_4 + 3H_2SO_4 = K_2SO_4 + 2MnSO_4 + 8H_2O + 5O_2$. Consequently, if an aqueous solution, containing a known amount of potassium permanganate, be run from a burette into a known volume of hydrogen peroxide until the pink colour of the permanganate is no longer discharged, it follows, from the equation, that every two molecules of $KMnO_4$ correspond with five molecules of H_2O_2 ; or 2×158.02 (the molecular weight of $KMnO_4$) grams of potassium permanganate correspond with 5×34.02 (the molecular weight of H_2O_2) gram of hydrogen peroxide; otherwise expressed, 1 gram of potassium permanganate represents 0.5382 gram of hydrogen peroxide.

EXAMPLE.—45 c.c. of a standard solution of potassium permanganate containing 20 grams of $KMnO_4$ per litre were decolorized by 25 c.c. of a solution of hydrogen peroxide. What amount of H_2O_2 is present in a litre of the hydrogen peroxide? Here 1000 c.c. of the standard solution contain 20 grams of $KMnO_4$; hence, 1 c.c. contains 0.02 gram; or 45 c.c. contain 0.9 gram. But from the equation, 1 gram of $KMnO_4$ represents 0.5382 gram of H_2O_2 ; hence, 25 c.c. of hydrogen peroxide has $0.5382 \times 0.9 = 0.4844$ gram of H_2O_2 . Hence a litre will have 19.4 grams of H_2O_2 .

Qualitative tests for the detection of hydrogen peroxide.—Special attention may be called to the titanium sulphate test, the chromic acid

test, the potassium iodide test; and the methods for distinguishing ozone from hydrogen peroxide. If a mixture of ozone, hydrogen peroxide, chlorine, and nitrogen peroxide be passed through a solution of potassium permanganate ozone alone escapes and will react with potassium iodide test-paper. A comparison of some reactions of ozone, hydrogen peroxide, chlorine, and nitrogen peroxide is given in Table VIII.

TABLE VIII.—DISTINGUISHING TESTS FOR OZONE AND HYDROGEN PEROXIDE.

Reagent.	Ozone.	Hydrogen peroxide.	Chlorine.	Nitrogen peroxide.
KI and starch .	Blue	Blue	Blue	Blue
Indigo solution .	Bleached	Bleached	Bleached	Bleached
Litmus solution .	Bleached	Bleached	Bleached	Bleached
Clean silver foil .	Blackened	Nil	White film	No change
Mercury bead .	Trail on glass	Nil	White	No change
CrO ₃ and ether .	Nil	Blue	Nil	Nil
Titanic acid .	Nil	Yellow	Nil	Nil
Tetramethyl base .	Violet	Nil	Blue	Yellow to brown
KMnO ₄ solution .	Unchanged	Decolorized	Decolorized	Decolorized

§ 6. Hydrogen Peroxide—Composition and Constitution.

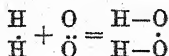
In spite of the fact that the composition and molecular weight of hydrogen peroxide have been determined; in spite of the simplicity of the resulting formula— H_2O_2 ; and in spite of the many investigations which have been made to determine the disposition of its constituent atoms, the constitution of this compound has not yet been settled by unequivocal experiments.

1. Empirical formula.—J. Thénard (1818) introduced a weighed amount of the peroxide in a small vial into a graduated cylinder over mercury. The vial was broken and its contents decomposed either by introducing manganese dioxide, or by heat. 17 parts of hydrogen peroxide by weight gave nearly 8 parts by weight of oxygen, and $17 - 8 = 9$ parts by weight of water. Otherwise expressed, 34 parts of hydrogen peroxide give 18 parts of water and 16 parts of oxygen. Hence the peroxide contains hydrogen and oxygen in the proportion of 2 atoms of hydrogen : 2 atoms of oxygen. The simplest formula for hydrogen peroxide is therefore HO. There is here nothing to show whether HO or some multiple of HO, say H_nO_n , is the proper formula for the compound, since the latter has the same percentage composition as the former.

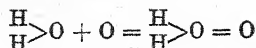
2. Molecular formula.—The instability of hydrogen peroxide prevents a determination of its vapour density being made in the regular manner. The molecular weight can be determined by the “freezing point method” which will be described later. The result is nearly 34. This agrees with the formula H_2O_2 —the generally accepted value.

3. Constitutional or graphic formula.—The formation of hydrogen peroxide by the action of oxygen on hydrogen, as the latter is liberated during the electrolysis of water, might lead to the supposition that its formula is $\text{HO} - \text{OH}$ (or $\text{HO} \equiv \text{OH}$). This is supported by the fact that

hydrogen peroxide is formed in reactions involving the reduction of oxygen:



and not by the oxidation of water



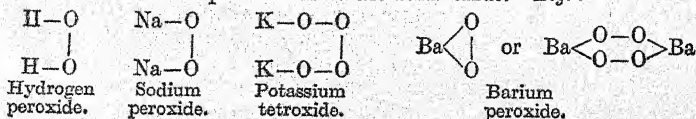
as would be expected if the formula were $\text{H}_2 = \text{O} = \text{O}$. Hence, by analogy it might be argued that the peroxides Na_2O_2 , etc., are similarly constituted, $\text{NaO} - \text{ONa}$. Two different lines of argument lead to two different conclusions, for J. W. Brühl showed that the refractive index of hydrogen peroxide is not characteristic of substances containing bivalent oxygen, and he infers that the oxygen is quadrivalent, $\text{H} - \text{O} \equiv \text{O} - \text{H}$. The two formulæ $\text{HO} - \text{OH}$ and $\text{HO} \equiv \text{OH}$ are similar in type, and at present it does not matter which is employed.

Purified hydrogen peroxide behaves like a mono- and dibasic acid in that (1) it reddens blue litmus; and (2) displaces the acid from sodium silicate, borate, metaphosphate, and sulphide, and from potassium ferro- and ferricyanides. Water liberates hydrogen peroxide from peroxides, and hence also water has been called an acid—*water acid*. (3) When solutions of the alkali hydroxides are treated with hydrogen peroxide, hydrated peroxides are produced—with barium hydroxide, $\text{BaO}_3 \cdot 8\text{H}_2\text{O}$ is precipitated. Conversely, when the peroxides are dissolved in dilute acids, the hydrogen peroxide is set free. In both cases, hydrogen peroxide behaves like an acid and the peroxides can accordingly be regarded as salts of hydrogen peroxide just as the nitrates are salts of hydrogen nitrate or nitric acid. (4) Hydrogen peroxide can be more easily extracted from its ethereal solution by alkali-lye than by water; (5) it partly neutralizes alkalis by entering into chemical union with them; (6) the peroxides are precipitated from solution of their salts—*e.g.* lead acetate and hydrogen peroxide give acetic acid and lead dioxide, and likewise also with barium salts, while lead dioxide is decomposed by an excess of hydrogen peroxide, barium dioxide forms *barium hyperoxide*, BaH_2O_4 , which breaks down when treated with more hydrogen peroxide into barium dioxide. Barium hyperoxide is regarded as a salt of the monobasic acid H_2O_3 , or as an acid salt of the dibasic acid, H_2O_2 . In either case, the graphic formula may be



D. I. Mendeléeff subdivides the peroxides into two classes depending upon the valency of the element united to the oxygen atoms:—

1. **Superoxides or true peroxides.**—Those peroxides in which the oxygen atom or atoms, over and above those required to form the basic oxide, are *simply linked* to the metal and to the other oxygen atoms, so as to form a chain. The valency of the metal is the same in the peroxide as in the basic oxide. *E.g.*:



The superoxides are often called *peroxidates*, being regarded as "salts" of hydrogen peroxide (p. 233).

2. **Polyoxides, dioxides, or pseudo-peroxides.**—Those peroxides in which the oxygen atom or atoms, over and above those required to form the basic oxide, are *doubly linked* to the metal so that the valency of the metal in the peroxide is *greater* than the valency of the metal in the basic oxide. *E.g.* :



Ozone.



Lead peroxide.



Manganese peroxide.

These oxides possess feeble basic or feeble acidic properties—possibly both.

The superoxides which yield hydrogen peroxide when treated with water or a dilute acid are probably constituted like hydrogen peroxide. Thus sodium peroxide with hydrochloric acid gives hydrogen peroxide; and potassium tetroxide, which gives oxygen and hydrogen peroxide, is probably constituted on the same plan. The polyoxides are not usually attacked by dilute acids. Both types with concentrated sulphuric acid evolve oxygen. The mechanism of the reaction is probably different in the two cases. With the superoxides, hydrogen peroxide is probably formed as an intermediate product, thus, with barium dioxide: $\text{BaO}_2 + \text{H}_2\text{SO}_4 = \text{BaSO}_4 + \text{H}_2\text{O}_2$; followed by $2\text{H}_2\text{O}_2 = 2\text{H}_2\text{O} + \text{O}_2$. With manganese dioxide: $2\text{MnO}_2 + 2\text{H}_2\text{SO}_4 = 2\text{MnSO}_4 + 2\text{H}_2\text{O} + \text{O}_2$. Similarly with hydrochloric acid, both give chlorine, but with the superoxides hydrogen peroxide is first formed, and this reacts with the excess of acid forming chlorine: $2\text{HCl} + \text{H}_2\text{O}_2 = 2\text{H}_2\text{O} + \text{Cl}_2$; with the polyoxides, on the other hand, an intermediate perchloride can often be detected—with manganese dioxide, probably MnCl_3 ; and with lead peroxide, lead tetrachloride, PbCl_4 is formed.¹

§ 7. Thermochemistry.

Sine igni nihil operamur.—C. GLASER.

Heat and cold are Nature's two hands by which she chiefly worketh.—F. BACON.

It was convenient, in a preceding paragraph, to resolve matter, as we know it, into two abstractions—matter and energy. Neither exists alone. We have no acquaintance with the one apart from the other. Isolated, matter and energy are pure abstractions. Each one completes and presupposes the other. The element phosphorus, for instance, must be regarded as a form of matter which is always associated with a certain amount of available energy, because it is able to do chemical work, and we cannot conceive of energy coming from nothing. We cannot answer: How much energy is associated with the phosphorus? The actual

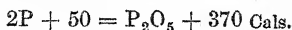
¹ Attempts have been made to show that two of the best known "polyoxides," PbO_2 and MnO_2 , are differently constituted because lead dioxide when exposed to sulphurous acid, H_2SO_3 , furnishes lead sulphate, PbSO_4 , while manganese dioxide furnishes manganous dithionate, MnS_2O_6 . It is more probable that the action in both cases is similar, manganese dioxide forming the normal sulphite: $\text{Mn}(\text{SO}_3)_2$; and lead dioxide, the basic sulphite: $\text{PbO} \cdot \text{SO}_3$. Both salts then undergo internal rearrangement, the former producing a dithionate, and the latter a normal sulphate.

amount available possibly depends upon the nature of the substance with which it is brought in contact. Similarly with oxygen. When these two elements—oxygen and phosphorus—are brought in contact, under the right conditions for the degradation of energy, chemical action sets in, and the chemical energy is degraded or transformed into heat or light. The resulting compound—phosphorus pentoxide—still contains some chemical energy, for if it be mixed with water, a great amount of heat is developed, chemical energy is degraded, and phosphoric acid results. The phosphoric acid still contains chemical energy because more energy is degraded in the form of heat when the phosphoric acid is brought into contact with sodium hydroxide. Every chemical reaction involves a change both in the form of the matter and in the form of the energy of the system. Modern chemistry would make Glaser's motto read: "Without energy nothing can change." What is generally understood by "descriptive" or "material" chemistry deals with the former, not with the latter. Chemistry proper is essentially concerned with both energy and matter, and hence it can no longer leave the development of the energy concept, as was formerly done, to physics.

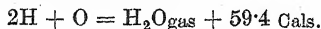
Law of Lavoisier and Laplace.—The free or available chemical energy of different substances is usually degraded in the form of heat during chemical action. The system gets hotter because heat is evolved by the reacting substances—such reactions are said to be *exothermal* or *thermo-positive reactions* in contrast with *endothermal* or *thermo-negative reactions* which absorb heat and thus cause the system to become cooler. That branch of chemistry which deals with the relation between thermal and chemical energy is called **thermochemistry**. Experiment shows that a definite chemical process generates a definite quantity of heat. Every compound has a definite heat of formation, which is numerically equal to the heat required for the decomposition of the compound back into its elements, but of opposite sign. "Action and reaction," said Newton, "are equal and opposite;" this is obviously a corollary of the law of conservation of energy. If it were not so, heat would be gained or lost when a compound is formed and then decomposed back into its original constituents. Such a result is at variance with the principle of the persistence or conservation of energy. The fact that every compound has a definite heat of formation which is numerically equivalent to its heat of decomposition but of opposite sign, is sometimes called the law of Lavoisier and Laplace, because A. L. Lavoisier and P. S. de Laplace first pointed out this generalization between 1780-84. Consequently, under a given set of conditions each atom and each molecule is charged with an amount of energy which is as definite as the quantity of matter itself. In every chemical reaction there is a definite change in the amount of energy so that after the reaction the products are associated with less available energy than was possessed by the initial products; part of the energy is frittered away as heat or electrical energy. Hence, if the process is to be reversed, and the original products restored, a supply of energy is needed to do work against the affinities of the products of the reaction, and to re-charge the original components with the required quota of energy.

The symbols used in thermochemistry.—It will be remembered that in physics, the unit of heat is the *calorie*, and a calorie represents the amount of heat required to raise the temperature of one gram of water through

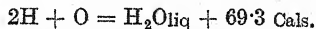
1° C. Consequently, 100 cals. will raise the temperature of 100 grams of water 1°, or of 1 gram of water 100°. In chemistry, it is convenient to represent the thermal value of a reaction by reference to the formula weights of the substances concerned in the reaction. Thus the heat of formation of phosphorus pentoxide is 370,000 cals. This means that 370,000 cals. are generated when 142 grams of phosphorus pentoxide are formed by burning 62 grams of phosphorus in oxygen; or 62 grams of phosphorus burning in oxygen will give sufficient heat to raise the temperature of 370,000 grams of water 1°. To avoid dealing with large numbers it will be more convenient to consider a Calorie as the amount of heat required to raise the temperature of 1000 grams of water 1° C. This is the so-called "kilogram-calorie," or the "big calorie," and it is here written with a capital "C." Thus "cal." refers to the gram-calorie. "Cal." refers to the kilogram-calorie. Hence the energy degraded in the form of heat when phosphorus burns in oxygen is equivalent to 370 Cals. This is represented in symbols:



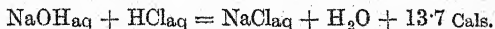
If the states of aggregation of the initial and final products of the reaction are not self-evident, the states of aggregation must be represented in the equation, otherwise latent heats of fusion or vaporization may lead to ambiguity.¹ Thus, the observed value



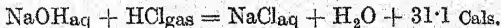
means that the union of 2 grams of hydrogen with 16 grams of oxygen is attended by the evolution of 59.4 big calories when the water produced is in the form of steam; if the steam is condensed to a liquid, the observed value is



The extra 9.9 Cals. represents the heat given out when 18 grams of steam are condensed to a liquid (*vide infra*). For solid water, another 1.4 Cals. are required to allow for the heat of solidification of 18 grms. of water. Again, if the reacting substances are in aqueous solution, a certain amount of heat may, or may not, be dissipated in the act of solution. For example, 13.7 Cals. are evolved when a dilute solution of sodium hydroxide is mixed with a dilute solution of hydrochloric acid. The dilute solution is represented by a suffix "aq." Thus,



If the sodium chloride were prepared by passing hydrogen chloride gas into a dilute solution of sodium hydroxide, more heat is evolved, because 17.4 Cals. are evolved when 36.4 grams of hydrogen chloride are dissolved in water:



A considerable amount of heat is absorbed during the formation of ozone— $3O_2 = 2O_3 - 68.2 \text{ Cals.}$ Hence, the molecules of ozone are charged with energy at a higher potential than those of oxygen, and when the reverse change occurs this energy is degraded in the form of heat. The fact that

¹ Some represent gases by means of italics, solids by clarendon type, and liquids by ordinary type. Some would represent the first equation $[2P, 5O] = 370 \text{ Cals.}$

the molecules of ozone are charged with a large proportion of energy is taken to explain why ozone decomposes so readily into oxygen.

Some disturbing effects.—Again, the chemical and physical condition of the reacting substances must be taken into consideration. The heats of combination of hydrogen in oxygen and in ozone would not be the same because of the reaction: $2\text{O}_3 = 3\text{O}_2 + 68.2 \text{ Cals.}$ Allowance would have to be made for this extra energy associated with ozone.

It is necessary to distinguish clearly between the observed heat changes and the real heat changes due to the degradation of chemical energy as heat. The observed thermal value of a reaction may be greater or less than that which corresponds with the chemical energy actually degraded during a given chemical reaction. For example, suppose that we start with a mixture of two volumes of hydrogen and one volume of oxygen, and finish with liquid water, there is a tremendous contraction in volume. This contraction occurs under atmospheric pressure (76 cm. mercury). Hence the atmosphere does work *on* the system, and that work appears as heat which raises the temperature of the system, and makes the observed heat of combination appear greater than it really is. Since $2\text{H} + \text{O} = \text{H}_2\text{O}_{\text{gas}} + 59.4 \text{ Cals. (observed)}$, and the observed latent heat $\text{H}_2\text{O}_{\text{gas}} = \text{H}_2\text{O}_{\text{liq}} + 9.9 \text{ Cals.}$; or $2\text{H} + \text{O} = \text{H}_2\text{O}_{\text{liq}} + 69.3 \text{ Cals.}$ The thermal value of the change of steam to water is partly due to the work of contraction. The work can easily be calculated,¹ and it is equivalent to 0.9 Cal. Hence

Apparent energy degraded in reaction	69.3 Cals.
Energy due to contraction	0.9 Cal.
Energy actually due to reaction	68.4 Cals.

Heat may also appear to be generated during a chemical reaction which is partly due to differences in the specific heats of the initial and final products of the reaction. If the latter be less than the former, some of the heat generated will be due to this fact, and not to the degradation of chemical energy.

§ 8. The Principle of Maximum Work.

The heat developed during a reaction represents a certain amount of potential energy which was associated with the atoms in some way; we are therefore much tempted to generalize, as Thomsen did in 1853, and assume that the thermal value of a reaction is a measure of chemical affinity between the reacting substances, and that every chemical change which can take place without the aid of external energy will be accompanied by the evolution of heat; or as M. Berthelot expressed it in 1869: Every chemical change which takes place without the aid of external energy tends to the production of that which is accompanied by the development of the maximum amount of heat. This is the so-called

¹ One gram-molecule of steam occupies 22.3 litres. The gases from which the steam was formed occupied $1\frac{1}{2}$ times this volume, i.e. 33.45 litres. A column of mercury, 1 sq. cm. sectional area and 76 cm. long weighs $76 \times 13.59 = 1033$ grams (since the specific gravity is 13.59). This pressure exerted along a path of 33.45 cm. will be $33.45 \times 1033 \text{ gram cms.}$, or, $33.45 \times 1.033 \text{ kilogram cms.}$ But 42.65 kilogram cms. are equivalent to one calorie. Hence, $33.45 \times 1033 \div 42.65 = 900 \text{ cals. or } 0.9 \text{ Cal.}$

principle of maximum work. Otherwise expressed, elements with the stronger affinity for one another liberate most heat when they react chemically. Hence also reactions which proceed spontaneously, when once they have been started, liberate some form of energy, generally heat, during the progress of the reaction.

Objections.—There are some objections to this generalization which cannot therefore be true in its present form:

(1) The principle assumes that reactions proceed completely to an end. In balanced reactions (p. 119), the reaction may be exothermal in one direction, and endothermal in another. According to the principle of maximum work, the exothermal reaction ought to go completely to an end. Hence the principle is not in agreement with facts.

(2) Several spontaneous reactions are known to be accompanied by an absorption of heat.

(3) Many systems require a preliminary impulse (p. 136) to start the reaction, and hence it would be necessary to introduce a clause to provide for this phenomenon.

The principle of maximum work must therefore be either amended or abandoned—every change which takes place without the aid of external energy must do work; and a system which cannot do work is incapable of spontaneous change and is in stable equilibrium. Further investigations have shown that it is not at all improbable that all possible chemical and physical reactions will be exothermal and complete at absolute zero, -273° ; and consequently, the principle of maximum work will probably apply at that temperature. At ordinary temperatures, the principle is only roughly applicable.

§ 9. The Principle of Reversibility.

The most stable compounds are usually, but not always, those with the greatest heat of formation. In a general way, the higher the temperature the less the stability of exothermal compounds; and conversely, endothermal compounds generally become more stable at higher temperatures because an absorption of heat is necessary for their formation. Here, then, we have another illustration of "the principle of reversibility" previously discussed, p. 38. A compound formed by an evolution of heat is decomposed by the addition of heat; and a compound formed by an absorption of heat is decomposed by the withdrawal of heat. Ozone is an endothermal compound; the equilibrium conditions at different temperatures in the presence of oxygen are:

	0°	100°	500°	1000°	2000°	3000°
Ozone	9.5×10^{-16}	3.3×10^{-11}	9.6×10^{-5}	2.2×10^{-2}	0.9	3.6 per cent.

Quite an appreciable amount of ozone may appear to be in equilibrium with oxygen at ordinary temperatures, but the amount for stable equilibrium is inappreciable. W. Ostwald (1891) has said:

It is generally believed that at a high temperature, such as that which exists in the electric arc, and in the sun's atmosphere, all compounds must be dissociated into their elements. This view is certainly not justified. On the contrary, what we actually know about the stability of compounds is that all compounds which are formed with an absorption of heat become more stable with rising temperatures, and *vice versa*. Owing to the fact that the majority of compounds known to us are formed from their elements with the evolution of heat, and in

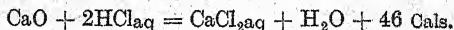
consequence, become more unstable as the temperature rises. it has been concluded that this is generally the case. But if we remember that acetylene and cyanogen—two compounds formed with the absorption of heat—are readily formed in quantity at the high temperature of the blast furnace, and in the arc light, we see the possibility that spectra occurring at high temperatures may belong to compounds which exist only at elevated temperatures.

As a matter of fact, some endothermal compounds become exothermal at higher temperatures—*e.g.* the formation of hydrogen iodide, hydrogen sulphide. Conversely, some exothermal compounds become endothermal at higher temperatures—*e.g.* silicon hexachloride. Consequently, there may be a reversion in the thermal value of some chemical processes when the temperature is augmented. The result is that a compound may be unstable at low temperatures and stable at higher temperatures; or conversely, stable at a low temperature and unstable at a higher temperature. Hydrogen peroxide and ozone are examples of the former; water an example of the latter.

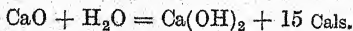
§ 10. Hess' Law.

Each element as well as each compound embodies a distinct and invariable amount of energy as well as a distinct and invariable quantity of matter, and thus energy is as constitutive and essential as a part of the existence of such element or compound.—J. B. STALLO.

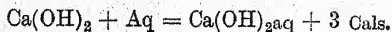
G. H. Hess (1840) measured the heat developed during the formation of a compound made in several different ways and came to the conclusion that the amount of heat evolved during the formation of a given compound is the same whether the compound is formed directly all at once or slowly in a series of intermediate stages. This is called Hess' law. The principle may be illustrated by making calcium chloride by the action of quicklime on dilute hydrochloric acid. It is found that :



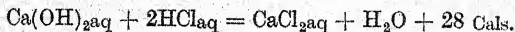
Instead of this, first slake the quicklime, and



Dissolve the calcium hydroxide in water, and



Mix the lime water with dilute hydrochloric acid, and



These three steps in the formation of the solution of calcium chloride give a total $28 + 3 + 15 = 46$ Cals. as the heat of formation. The same result was obtained by the direct action of the dilute acid on quicklime. A number of experiments made on similar lines have shown that (1) The heat of formation of a compound is independent of its mode of formation; and (2) the thermal value of a reaction is independent of the time occupied by the change.

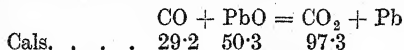
This result is but a particular application of the law of persistence of energy, and it may be expressed by saying that the change of energy of a system in passing from one state to another depends upon the initial and final states of the system, and not on the intermediate states. Starting with given raw materials, suppose that it were possible to make a compound by two different processes so that the total heats of formation of

the compound were different, it follows that a decomposition of the compound formed by one of the processes (Lavoisier and Laplace's law) could lead to the creation or destruction of energy.

It follows from Hess' law that if the heat of formation of carbon dioxide be: $C + 2O = CO_2 + 94.3$ Cals.; and $CO + O = CO_2 + 68$ Cals., we have

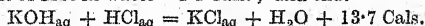
$$(C + 2O) - (CO + O) = 94.3 - 68 \text{ Cals.}$$

Consequently, the heat of formation of carbon monoxide is: $C + O = CO + 26.3$ Cals. This illustrates the fact that the thermal value of a reaction is the sum of the heats of formation of the final products of the reaction less the heats of formation of the initial products of the reaction. This corollary to Hess' law is valuable because it enables the heat of formation of a compound from its elements to be computed when a direct determination is either impracticable or very difficult. Similarly the thermal value of a reaction can be deduced when the heats of formation of the different substances which take part in the reaction are known. Thus, writing the heats of formation of the substances concerned in the following reaction below their chemical symbols,

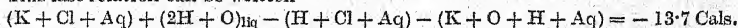


it follows that the heat of the reaction is $97.3 - (29.2 + 50.3) = 17.8$ Cals. It will not be forgotten that the energy of a reaction is not primarily inherent in any one of the reacting substances, but belongs to the system as a whole; this energy can be regarded as the sum of two or more constants which are peculiar to the respective elements involved in the change.

EXAMPLE.—It is required to compute the heat of formation of $K + Cl = KCl$, when it is known that the heat of formation of $K + O + H + Aq = KOH_{aq} + 116.5$ Cals.; $2H + O = H_2O_{liq} + 68.4$ Cals.; $H + Cl = HCl_{aq} + 39.3$ Cals.; heat of solution of KCl in water— 4.4 Cals.; and that



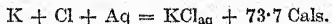
This last relation can be written



Consequently, after substituting the given data, we get

$$(K + Cl + Aq) + 68.4 - 39.3 - 116.5 = -13.7$$

Hence,



Subtract the heat of solution — 4.4 Cals., and we get $73.7 - (-4.4) = 78.1$ Cals. for the thermal value of the reaction $K + Cl = KCl$. It will be noticed that the solution of potassium chloride in water is an endothermal process, and hence the heat of formation of KCl_{aq} is less than the heat of formation of KCl .

There is a relatively large error in the determination of the heats of chemical reactions, and these errors may be magnified very much when the thermal value of a reaction is obtained indirectly by calculations based on Hess' law.

Questions.

1. Write an account of the chemistry of ozone. In what way is the composition of ozone deduced?—*St. Andrews Univ.*
2. Finely divided silver thrown into hydrogen dioxide occasions a sudden evolution of oxygen, but the metal is not oxidized. Silver oxide in like manner occasions a similar evolution, and metallic silver results. Explain these facts.—*Science and Art Dept.*
3. What are the chief properties of hydrogen peroxide? 20 c.c. of a solution

of this substance, after acidification with dilute sulphuric acid, reduced 24 c.c. of $\frac{1}{10}$ *N*-potassium permanganate. Calculate the percentage of hydrogen peroxide in the solution.—*St. Andrews Univ.*

4. What is meant by the terms "endothermic" and "exothermic" compounds? To what class do substances known as "explosives" belong? Account for the greater activity of ozone over that of oxygen on thermochemical grounds, and for its resolution into oxygen by compression.—*Science and Art Dept.*

5. What is a "silent discharge"? Make a sketch of a piece of apparatus suitable for submitting gases to such discharge, and indicate any chemical changes which can be produced by its means.—*London Univ.*

6. What is the heat of formation of zinc chloride, ZnCl_2 , in solution when the reaction, $\text{Zn} + 2\text{HCl}_{\text{aq}} = \text{ZnCl}_{2\text{aq}} + \text{H}_2 + 34.4 \text{ cal.}$, and the heat of formation of an aqueous solution of hydrogen chloride, is given by the equation: $\text{H}_2 + \text{Cl}_2 = 2\text{HCl} + 78.6 \text{ cal.}$?—*French Coll.*

7. Discuss the question of the valency of oxygen referring specially to the evidence of its character as a quadrivalent element.—*Board of Educ.*

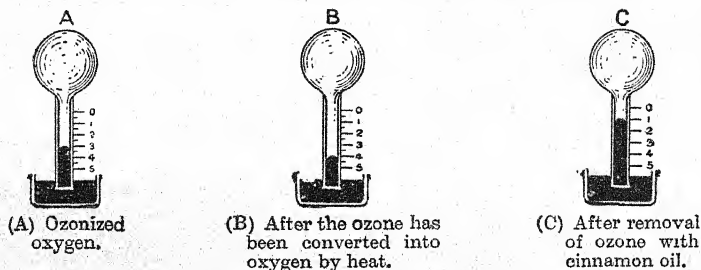


FIG. 88.—Soret's Experiments on Ozone.

8. Five cubic centimetres of a solution of hydrogen peroxide were mixed with an acidified solution of potassium permanganate and 20 c.c. of oxygen were evolved measured at N.T.P. From this result, calculate the per cent. of hydrogen peroxide present in the solution. How would such a solution be labelled in commerce?—*Board of Educ.*

9. When hydrogen is heated to 2100° , it has been stated that some of the diatomic molecules are decomposed $\text{H}_2 \rightleftharpoons 2\text{H}$; again ozone is formed when oxygen is subjected to an electric discharge: $3\text{O}_2 \rightleftharpoons 2\text{O}_3$. Are the substances symbolized by O_2 , H_2 , and O_3 elements? How do the facts just indicated affect the definition of an element? An element is a substance which cannot be separated into simpler substances by any known means?

10. How may a chemical equation be modified so as to represent not only a redistribution of matter but also a redistribution of energy? Explain what is meant by the heat of formation of a compound. From the following data find the heat of formation of acetylene, C_2H_2 : When 24 grms. of carbon, 2 grms. of hydrogen, and 26 grms. of acetylene are burnt separately in an excess of oxygen, 194,000, 68,000, and 310,000 calories are respectively evolved.—*Cape Univ.*

11. Find the amount of heat *K* libera ted in the reaction $\text{AlCl}_3 + 3\text{Na} = 3\text{NaCl} + \text{Al} + xK$. The heat of formation of aluminium chloride, AlCl_3 , is 1610K, and of sodium chloride, NaCl , 976K.—*Sydney Univ.*

12. What amount of heat is evolved when 46 grms. of metallic sodium react with an excess of water, given the heats of formation of water (H_2O) as 69 Cals., and of sodium hydroxide (NaOH), 112.5 Cals.?

13. What substances besides ozone have the power of turning iodized starch paper blue? How are they distinguished from ozone?—*Science and Art Dept.*

14. Write a concise historical account of the chemistry of ozone as developed by the researches of Schönbein, Andrews and Tait, Brodie, Soret, Hautefeuille and Chappius, Houzeau and others. How do you account for ozone being an explosive substance under certain conditions? Define the conditions under which it may be exploded.—*Science and Art Dept.*

15. Find the heat of formation of aldehyde, $\text{C}_2\text{H}_4\text{O}$, from its elements when aldehyde is (a) liquid, and (b) gaseous. Data: $\text{C}_2\text{H}_4\text{O} + 5\text{O} = 2\text{CO}_2 + 2\text{H}_2\text{O}$;

[$\text{C}_2\text{H}_4\text{O}$, 50] liquid 275.5 Cals.; [CO_2] 96.6 Cals.; [H_2O] liquid 68.4 Cals.; [H_2O] gaseous 58.7 Cals.; [$\text{C}_2\text{H}_4\text{O}$, 50] gaseous, 266.0 Cals.—*Owens College*.

16. How is hydrogen peroxide prepared? Mention any instances known to you of its action (i) as an oxidizing agent, and (ii) as a reducing agent. What is the evidence that its molecule contains twice as much oxygen as a molecule of water?—*London Univ.*

17. In Soret's experiments a flask *A*, Fig. 88, containing ozonized air was heated to decompose the ozone, *B*, Fig. 88, and another flask containing the same volume of ozonized air contracted 6.90 c.c. when treated with cinnamon oil, *C*, Fig. 88. Show that this agrees very nearly with the inference that 2 c.c. of ozone furnishes 3 c.c. of ordinary oxygen. In another experiment the contraction after treatment with cinnamon oil was 8.0 c.c., and with the same amount of ozonized air, the oxygen absorbed by treatment with potassium iodide without change of volume was 3.9 c.c. Show that this means that 8 c.c. of ozone is nearly equivalent to 8.0 + 3.9 c.c. of oxygen.

CHAPTER XII

OSMOTIC PRESSURE AND RELATED PHENOMENA

§ 1. Diffusion in Gases and in Liquids.

The act of dissolution is probably not due to chemical combination in the first place, but is probably analogous to the sublimation of a solid into a gas, and proceeds from the detachment of molecules from the surface of the solid, and their intermixture with those of the surrounding liquid. This is doubtless due to the impact of the moving molecules of the liquid.—
W. G. TILDEN.

LET a large crystal of a coloured salt—say copper sulphate—be placed at the bottom of a tall glass cylinder, and the remainder of the jar be filled with water. The coloured salt is chosen because the movements of the resulting solution can be readily seen. Let the jar stand where it will not be disturbed by evaporation, agitation, etc. The surface of separation between the solid and solvent will be gradually obliterated; in time, the coloured salt will diffuse uniformly throughout the whole body of liquid. The diffusion of the salt in the solvent seems to be analogous with the process of diffusion in gases. It is inferred that the molecules of the liquid are in perpetual motion in all directions; and that the protracted time occupied by the diffusion of the molecules of the dissolved salt in the liquid is due to the close packing of the molecules of the liquid. Consequently, the free progress of the molecules of the dissolved salt—p. 139—in the solvent is greatly impeded.

Just as the molecules of a gas in a closed vessel are disseminated in a relatively large space, so are the molecules of a solid in solution scattered in a relatively large volume of solvent. It is true that the molecules of the salt in solution could not occupy the space if the solvent were absent, otherwise the analogy between a substance dissolved in a solvent and a gas scattered in space would be very close. Arguments from analogy are notoriously treacherous; and whatever conclusions might be inferred from a closer study of the analogy between the process of solution and gaseous diffusion, the fact that the molecules of the dissolved substance are co-mingled with the solvent, and that the molecules of the gas are not associated with such an agent, must be constantly borne in mind. As G. F. Fitzgerald has said: "The dynamical condition of molecules in solution is essentially and utterly different from that of the molecules of a gas."

§ 2. Solution Pressure—Osmotic Pressure.

Just as a small quantity of water is able to dissolve a quantity of salt which can diffuse itself through a large quantity of water, so a quantity of air which can expand and diffuse itself through a large space may be contained within a small compass.—R. HOOKE (1664).

If the diffusion of gases be resisted by placing a permeable partition between two gases, a pressure will be exerted upon the partition, as was exemplified in the experiments on p. 129. It is easy to show that the particles of a dissolved substance exert a similar pressure when a partition is placed between the solution and solvent so that the partition offers no obstacle to the free circulation of the molecules of the solvent, but resists the free passage of the molecules of the dissolved substance.

A piece of wet bladder is stretched and wired over the head of a wide thistle-headed funnel with a stem about 10 cm. long. When nearly dry, the bladder is removed and the hot funnel is smeared about the rim with marine glue. The bladder is immediately wired securely in position. The thistle-headed funnel is nearly filled with a concentrated solution of cane sugar and joined by means of pressure tubing or a rubber stopper with a piece of capillary tubing of $\frac{1}{8}$ mm. bore bent S-shaped as indicated in Fig. 89. The funnel is immersed in a jar of water. The level of the index of coloured water in the capillary tube is marked with gummed paper, and the apparatus is allowed to stand over night. In the morning the liquid in the capillary will have risen about 10 cm. Water has obviously passed from the beaker through the membrane into the sugar solution.



FIG. 89.—Illustration of Osmotic Pressure.

The passage of water through a membrane in this manner is called osmosis—from the Greek *ὀσμός* (osmos), a push. If the osmosis be inwards, towards the solution, it is called endosmosis; if outwards, exosmosis. The membrane permeable to the solvent, impermeable to the dissolved substance, is called a semi-permeable membrane. The extra pressure exerted upon the membrane by the sugar solution was styled, by W. Pfeffer (1877), "the osmotic pressure of the sugar solution." Solutions with the same osmotic pressure are said to be *is-osmotic* or *isotonic*.

The action is curious. In the ordinary nature of things the sugar would diffuse into the solvent until the whole system had one uniform concentration. The membrane prevents this. If the sugar cannot get to the solvent, the solvent goes to the sugar—a case of Mohamet and the mountain. Molecules of sugar and molecules of water attempt to pass through the membrane; the way is open for the molecules of water, but not for the molecules of sugar. Water can pass freely both ways. The extra pressure on the solution side of the membrane—the solution pressure—is supposed to be due to the bombarding of the membrane by the molecules of sugar. Equilibrium occurs when the number of molecules of water passing downwards through the membrane is equal to the number passing in the opposite direction. The resulting pressure is the solution pressure or the osmotic pressure of the solution.

Let us be perfectly clear about this, or we may be led into error. The *fact* observed is that the osmotic pressure is the excess of the pressure on the solution side of a semipermeable membrane over the pressure on the solvent side. The *hypothesis* here suggested—often styled van't Hoff's hypothesis—is that this pressure is due to the bombarding of the semipermeable membrane by the dissolved molecules trying to diffuse into the solvent and make solvent and solution one uniform concentration. Harsh experience alone can shatter or establish this revolutionary hypothesis.

Imagine the experiment arranged a little differently. Suppose the aqueous solution of sugar in the lower part of a cylinder, Fig. 90, to be separated from the pure solvent in the upper part of the cylinder by a semipermeable membrane *A*, so fitted that it can slide freely up and down the cylinder. The upward osmotic pressure of the solution will naturally force the piston upwards, and a weight *P*, equivalent to the osmotic pressure of the solution, will be required to keep the semipermeable membrane

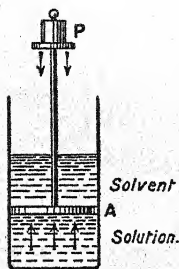


FIG. 90.—Osmotic Pressure.

in one fixed position. A. C. Brown has an interesting experiment illustrating this. A concentrated solution of calcium nitrate is saturated with phenol and the mixture poured into a tall narrow cylinder. The excess of phenol rises and floats upon the surface of the calcium nitrate solution. The phenol should not be in larger excess than is required to give a layer a few millimetres thick. Distilled water, saturated with phenol is carefully poured above the two layers of liquid in the cylinder. The water floats on the surface of the phenol. The water on both sides of the phenol can traverse the partition of phenol, but the calcium nitrate cannot pass through. Hence the layer of phenol is a semipermeable membrane. Mark the level of the layer of phenol in the cylinder by means of a piece of gummed paper. If the upward motion of the layer of phenol be marked from day to day, it will be found to rise higher and higher, and finally surmount the rest of the liquid in the cylinder.

§ 3. The Measurement of Osmotic Pressure.

The basis of all knowledge is experiment; the very essence of experiments is exactness; and exactness can be obtained only by precise measurements.
—J. T. SPRAGUE.

Animal membranes are objectionable when exact measurements are required because, to a certain extent, the results depend upon the nature of the membrane; the membrane is not strong enough to withstand the great pressures developed by osmosis; and, most serious of all, the membrane is not quite semipermeable, an appreciable amount of, say, sugar does actually pass through the membrane. It would therefore be as profitable to measure the pressure of a gas in a leaking vessel as to try to measure the osmotic pressure of a solution with a membrane which allows part of the dissolved substance to pass through. We therefore fall back on artificially prepared membranes. No artificial membrane has been

so successful as a film of copper ferrocyanide deposited ¹ between the inner and outer walls of a "porous pot," and illustrated by the sketch of a broken pot, *A*, Fig. 91. The porous pot with its semipermeable membrane *A* is fitted with a suitable manometer (Fig. 92) to indicate the pressure. W. Pfeffer made some measurements with cells made in this manner, in 1877. The apparatus was immersed in a large bath of water to maintain the temperature constant during the experiment.

§ 4. Osmotic Pressure and the Gas Laws.

The development of the bombardment theory of osmotic pressure is largely based on the analogy between the processes of solution and vaporization. The process of solution can be compared with vaporization; the dissolved substance is distributed throughout the whole bulk of the solution, and, pursuing the analogy, the solvent thus plays the part of so much space. The vapour pressure of a solvent in space will thus be represented by the osmotic pressure of a substance in solution. The closed space over a liquid becomes saturated with vapour, so does a solvent in contact with the solute form a saturated solution. An increase in temperature augments the vapour pressure of a liquid, and likewise too the osmotic pressure of a solution.

1. The relation between osmotic pressure and the concentration of the solution—Boyle's law.—W. Pfeffer (1877) obtained some data with this apparatus which J. H. van't Hoff (1887) utilized, with remarkable cleverness, in developing what he called "the rôle of osmotic pressure in the analogy between solutions and gases." The experimental data showed that the osmotic pressure is very nearly proportional to the concentration of the solution; otherwise expressed, the osmotic pressure appears to depend upon the degree of crowding of the molecules of the dissolved substance. Instead of repeating Pfeffer's measurements, a selection from some later determinations with solutions of glucose by H. N. Morse (1907) can be quoted (temperature nearly 0°), rounding off the decimals to the nearest tenth of a unit:

Concentration	0.1,	0.2,	0.3,	0.4,	0.5,	0.6,	1.0 ²
Osmotic pressure	2.4,	4.7,	7.0	9.3,	11.7,	14.1,	23.7 atm.
Equivalent gas pressure	2.2,	4.5,	6.7,	8.9,	11.1,	13.4,	22.3 atm.

¹ By steeping a clean porous pot in an aqueous solution of potassium ferrocyanide, rinsing in water, and then submerging the pot in an aqueous solution of copper sulphate.

² Morse's data are here rounded off, and only a few selected. One molecule weight of glucose expressed in grams will occupy 22.3 litres (p. 84). Hence 0.1 molecule will occupy 2.23 litres. By choosing the concentration so that

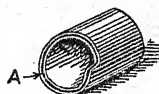


FIG. 91.—Semipermeable Membranes.

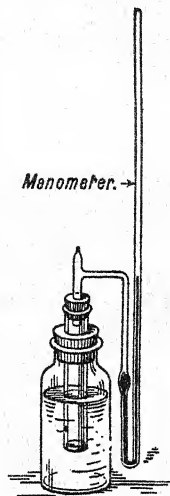


FIG. 92.—Measurement of Osmotic Pressure.

In dealing with the concentration of solutions, the same unit of comparison can be employed as in dealing with gases, viz., the molecular weight of the solute expressed in grams per 22.3 litres at n.p.t.

The "equivalent gas pressure" is here calculated on the assumption that a "sugar gas" obeying Boyle's law really exists. The results are plotted in Fig. 93. The deviation of the osmotic pressure curve from the dotted curve emphasizes the fact that the deviations of the osmotic from the equivalent "gas pressures" grow larger with increasing concentrations, and hence exact proportionality occurs only when the solutions are very dilute. For dilute solutions, the osmotic pressure is nearly proportional to the concentration, or, as W. Ostwald puts it, "the osmotic pressure of a sugar solution has the same value as the pressure the sugar would exert if it were contained, *as a gas*, in the volume occupied by the solution." This is another way of saying that the relation between the osmotic pressure of a solution and its concentration has the same form as Boyle's

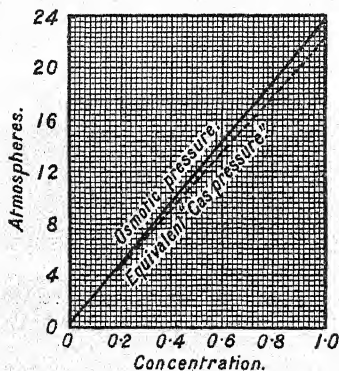


FIG. 93.—Osmotic Pressure and Concentration.

law for gases. The analogy does not work out so well for concentrated solutions as with dilute solutions—possibly owing to the disturbing effects of overcrowding produced by: (1) molecular attraction between the molecules of the dissolved substance; (2) and between the molecules of the solute and solvent; and (3) the volumes of the molecules themselves. The first and last of these effects for gases were discussed when dealing with Boyle's law for gases.

2. The relation between osmotic pressure and temperature—Charles' law.—Pfeffer's measurements on the influence of tempera-

ture also showed that the osmotic pressure is proportional to the absolute temperature, which means that the relation between the osmotic pressure and the temperature of a given solution has a formal analogy with Charles' law for gases. In illustration, some results by H. N. Morse (1911) for unit concentration may be quoted:

Temperature	0°	5°	10°	15°	20°	25°
Osmotic pressure	24.8,	25.3,	25.7,	26.2,	26.6,	27.0 atm.
Equivalent gas pressure	22.2,	23.0,	23.4,	23.8,	24.2,	24.5 atm.

The "equivalent gas pressure" is here calculated on the assumption that a "sugar gas" obeying Charles' law really exists. These numbers are plotted in Fig. 94, and the graphs show the proportionality between osmotic pressure, p , and temperature; $p/T = \text{constant}$. The space between the two curves represents the deviation of the observed osmotic

in Boyle's relation, $pv = \text{constant}$, a solution containing a molecular weight expressed in grams, per 22.3 litres, has a concentration of 22.3 units when $p = 1$, we get from Boyle's law $p \div C = 22.3$. The concentration, it will be remembered, is inversely proportional to the volume. Hence for a concentration 0.1, we get $p = 2.23$; for $C = 0.2$, $p = 4.46$, etc.

pressure from the pressure calculated on the assumption that the dissolved substance behaves as if it were a gas. One gram-molecule of a gas under standard conditions occupies nearly 22.3 litres; similarly one gram-molecule of a solute in 22.3 litres of solvent may be said to occupy 22.3 litres, and to exert an osmotic pressure equivalent to one atmosphere. If P denotes the osmotic pressure and V the volume, $PV = RT$. Hence, quantities proportional to the gram-molecular weights dissolved in the same quantities of solvent exert the same osmotic pressure, and if we were dealing with solutions containing n gram molecules of solute, $PV = nRT$.

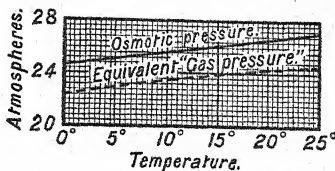


FIG. 94.—Osmotic Pressure and Temperature.

3. Avogadro's hypothesis applied to solutions.—By a comparison of the concentration of solutions at the same temperature and the same osmotic pressure, it has been inferred that they contain the same number of molecules of the dissolved substance per unit volume. The term "number of molecules" is used in the same sense that the term is used in stating Avogadro's hypothesis for gases: "Equal volumes of all gases at the same temperature and pressure contain the same number of molecules; or conversely, at any assigned temperature, the pressure of a gas depends on the number of molecules and not on their kind." Hence van't Hoff's hypothesis assumes that the osmotic pressure and related properties—vapour pressure, freezing point, and boiling point—of dilute solutions

(1) depend upon the number¹ of molecules of solute dissolved in unit volume of the solution, and (2) are independent of the chemical nature of the solvent and solute, and (3) of the relations between solvent and solute.

Equal volumes of solutions containing the same number of solute molecules have the same osmotic pressure. There is a striking resemblance between this assumption and Avogadro's hypothesis for gases, and it harmonizes with a number of facts. The principle can be applied to measure the molecular weight of substances in solution.

EXAMPLES.—(1) An aqueous solution of 1.0047 gram of orthoboric acid per litre at 0° has an osmotic pressure of 27.3 cm. of mercury. What is the molecular weight of the acid? Since 1 gram-molecule of a substance in the gaseous state occupies 22.3 litres at 0° and 760 mm., we have here to find what weight of substance will occupy 22.3 litres at 0° and 760 mm., given 1.0047 gram occupy 1 litre at 0° and 273 mm. pressure. Obviously, 1.0047 gram will occupy 0.361 litre at 0° and 760 mm.; and if 0.361 litre weighs 1.0047 gram, 22.3 litres will weigh 62 grams at the same temperature and pressure. Hence the molecular weight of the given acid is 62; this agrees with the formula $B(OH)_3$ for orthoboric acid.

(2) A 2 per cent. solution of cane sugar has an osmotic pressure of 1016 mm. at 15°; what is the molecular weight of cane sugar? 100 c.c. at 1016 mm. pressure becomes 0.126 litre at 760 mm. pressure and 0°, and 0.126 litre corresponds with 2 grams of cane sugar. Hence 22.3 litres will have 355 grams at the same temperature and pressure. The molecular weight of cane sugar is therefore 355. The true number is 342.

There are so many experimental difficulties involved in the direct

¹ COLLIGATIVE PROPERTIES.—In contradistinction to additive properties, colligative properties depend on the relative number of molecules present and not on the kind of molecules—e.g. the osmotic pressure and the freezing and boiling points of solutions.

measurement of osmotic pressures that the method is rarely if ever employed for molecular weight determinations.

§ 5. The Relation between the Vapour Pressure of a Solution and the Molecular Weight of the Solute.

M. Faraday knew, in 1822, that the vapour pressure of a solution is lower than the vapour pressure of the pure solvent; but A. Wüllner discovered the important fact experimentally, in 1858, that the lowering of the vapour pressure of a solution is proportional to the quantity of substance in solution provided that the dissolved substance is non-volatile. This is sometimes called Wüllner's law.

Suppose a solution *A*, Fig. 95, confined in a long-stemmed tube, as illustrated in the diagram, be separated by a semipermeable membrane *M* from the pure solvent. Let all be confined in a closed vessel.



FIG. 95.

pressure will force the solution to rise in the narrow tube to a height h , when the whole system is in equilibrium. Let p_s denote the vapour pressure of the solution in the narrow tube, and p the vapour pressure of the solvent in the outer vessel. The vapour pressure of the solution at the surface in the narrow tube must be equal to the vapour pressure of the solvent at the same level, otherwise distillation would take place either to or from the surface of the liquid in the narrow tube. In either case there would be a constant flow of liquid respectively to or from the vessel *A* through the semipermeable membrane in order that h may have a constant value. Otherwise expressed, perpetual motion would occur. By the "law of excluded perpetual motion" this is impossible, hence the vapour pressure of solution and solvent at the upper level of the solution in the narrow tube must be the same. The vapour pressure of the solvent

at the level *a* will be equal to the vapour pressure of the solvent at the lower level *b* less the pressure of a column of the height h or $p = p_s + w$. Since the height h is determined by the osmotic pressure, which in turn is determined by the concentration of the solution, there must be a simple proportionality between the osmotic pressure or concentration of the solution and the lowering of the vapour pressure ($p - p_s$). Just as the osmotic pressure of a dilute solution is proportional to the concentration of the dilute solution, so it can be proved that the vapour pressure is proportional to the osmotic pressure, and consequently, the relative lowering of the vapour pressure of a solvent by the addition of a foreign substance is proportional to the concentration. The phenomenon can be illustrated by introducing about 2 c.c. of water, 2 c.c. of a 2 per cent. solution of potassium iodide, and 2 c.c. of a 4 per cent. solution of the same salt into the Torricellian vacuum of three barometer tubes mounted within a hot jacket. The effect will be obvious from Fig. 96.

If w denotes the weight of substance in grams dissolved in 100 grams of solvent, and if p denotes the resultant lowering in the vapour pressure

of the solvent, it can be shown by an extension of the above reasoning that the molecular weight of the solute is :

$$\text{Molecular weight} = k \frac{w}{p}$$

where k is a constant whose numerical value depends upon the particular solvent used. The method for determining the molecular weight of a substance from direct measurements of the lowering of the vapour pressure is of great theoretical interest, but in practice the method is seldom employed, because some of the related properties of solutions are more amenable to accurate measurement.

Deliquescence.—If a soluble substance becomes moist by the condensation of moisture on its surface on exposure to the air the vapour pressure of the concentrated solution so formed is *less* than the vapour pressure of the moisture in the surrounding air. Hence more moisture condenses on the surface, and this continues until the vapour pressure of the solution is equal to the vapour pressure of the aqueous vapour in the atmosphere. Thus deliquescent substances not only become moist, but they attract so much moisture from the atmosphere that they dissolve in the water removed from the atmosphere; *e.g.* calcium chloride, potassium carbonate, sulphuric acid, etc.

The evaporation of solutions.—G. F. Fitzgerald (1896) has pointed out that the kinetic theory of evaporation describes the lowering of the vapour pressure of a solution in this manner: The presence of non-volatile molecules of the solute at the surface of the solution hinders the egress, but does not prevent, or possibly facilitates, the return of the volatile molecules (Fig. 56). The gas-analogy hypothesis of osmotic pressure assumes that the presence of a body in solution produces no effect or the same effect on the ingress or egress of the molecules of the solute, for the surface of a liquid with a non-volatile solute is a perfect semipermeable membrane—water molecules can pass through the surface freely, but the molecules of the solute cannot. It is a remarkable coincidence that with dilute solutions the osmotic pressure is roughly the same as that which would be produced by the molecules of the solute if it were in the gaseous

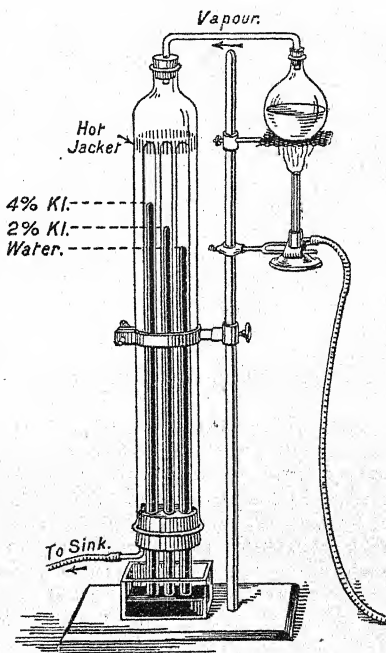


Fig. 96.—The Vapour Pressure of Solutions.

state, but, as previously indicated, the dynamical theory of the two must be intrinsically different.

§ 6. Callendar's Vapour Pressure Hypothesis of Osmotic Pressure.

The substitution of analogy for fact is the bane of chemical philosophy; the legitimate use of analogy is to connect facts together and to guide to new experiments.—H. DAVY.

Vague similarities in certain properties are never sufficient to determine a person who earnestly seeks for the truth and is not shackled by hypotheses.—J. BERGMAN.

The gas-analogy hypothesis of osmotic pressure.—The "laws" associated with the names of Boyle, Charles, Dalton, and Graham, and the hypothesis of Avogadro, are but a few of the many striking analogies subsisting between the behaviour of gases confined in a given space, and substances in dilute solution. We know enough about nature to believe that if two things are exactly alike, they will behave alike under the same circumstances; but when the things compared are not quite similar, we must be prepared for discrepancies. Analogy is not proof. Had Isaac Newton measured the refractory power of native cadmium sulphide—greenockite—he would no doubt have said: "greenockite is probably an unctuous substance coagulated," and he would have been wrong. As it happened, this prognostication turned out all right with the diamond. The hypothesis that the osmotic pressure of a dilute solution is produced by the bombardment of the semipermeable membrane by the dissolved molecules gives a very plausible interpretation of the analogy between the behaviour of dissolved molecules and the molecules of a gas brought out by J. H. van't Hoff, but, *in theory*, makes so little of the part undoubtedly played by the solvent, and, *in practice*, the analogy appears to break down so completely with more concentrated solutions that a number of rival hypotheses have been advanced to explain the phenomena. The principle of exhaustion, indicated on p. 20, compels us to investigate other hypotheses. H. L. Callendar's vapour pressure hypothesis (1909) is one of the most satisfactory of the purely physical explanations of osmotic pressure, and it is superior, in many respects, to the gas-analogy hypothesis. Callendar's hypothesis has been tested with somewhat concentrated solutions, and wherever data are available it has been eminently successful.

The vapour pressure of a liquid under pressure.—Experiment shows that the maximum vapour pressure of a solution can be altered in three ways: (1) by altering the temperature (p. 187); (2) by varying the concentration of the solution (Fig. 96); and (3) by altering the pressure under which the liquid itself is confined. The effect of pressure on the freezing point of water (ON, Fig. 61) is an application of the third principle.

The student might very properly raise the objection to the third method of altering the vapour pressure of a liquid; it has been shown, p. 187, to be impossible to raise the pressure on a saturated vapour without causing it to liquefy. If a vertical cylinder, provided with a piston, contains nothing but water-liquid and vapour, it is quite true that the descent of the piston will result in the condensation of water vapour until all the vapour is liquefied, and as long as water vapour is present the vapour pressure remains constant. On the contrary, if air as well as water vapour be present, it is easy to see that the volume of the air decreases, or the pressure of the air on the surface of the liquid increases during the

descent of the piston. The water vapour still supports its own share of the total pressure up to its maximum vapour pressure, and water vapour not quite so much as before will condense, consequently the liquid under a considerable external pressure can exert a greater vapour pressure than the maximum vapour pressure under atmospheric pressure.

The relation between vapour pressure and osmotic pressure.—It has been proved experimentally that the maximum vapour pressure of a solution under very great pressures is rather greater than the maximum vapour pressure of the same solution under atmospheric pressures, see the curve *ON*, Fig. 61. Again, the vapour pressure of a solution is *less* than the vapour pressure of the pure solvent, Fig. 61. Consequently, if the pressure on a solution be sufficiently augmented, the pressure of its vapour can be made equal to the vapour pressure of the pure solvent under atmospheric pressure. This is the condition necessary in order that solution and solvent can exist side by side in equilibrium. If the vapour pressure of the solution were less than that of the pure solvent, the system would not be in equilibrium, because vapour would distil from the solvent into the solution until the vapour pressure of both were the same. Conversely, when a solution under its own osmotic pressure and the pure solvent are in equilibrium, it follows that their vapour pressures must be equal. Hence, according to Callendar: The osmotic pressure of a solution represents the external pressure which must be applied in order to make its vapour pressure equal to that of the pure solvent. With this hypothesis, Callendar has calculated the osmotic pressures of sugar solutions of different concentration from published vapour pressure data, and the results are in close agreement with observation:

Concentration	180,	300,	420,	540 grams per litre.
Observed osmotic pressure . .	14.6,	26.8,	44.0,	67.5 atmospheres.
Calculated osmotic pressure . .	14.1,	26.8,	43.7,	67.6 atmospheres.

Hence it is inferred that osmotic equilibrium depends upon the equality of the vapour pressure of the solution and of the pure solvent.

A semipermeable membrane may be likened to a partition pierced by a large number of minute capillary tubes; suppose that the capillary tubes are not wetted by either the solvent or solution, then neither the liquid solvent nor the solution can enter the capillaries,¹ although vapour can diffuse through the capillary tubes. But the vapour pressure of the solution on one side of one of the capillary tubes is less than the vapour pressure of the solvent on the other side; consequently, vapour will pass through the capillary and distil from the solvent to the solution. Hence the volume of the solution will increase, and if the solution be confined in a closed vessel, the pressure must rise and continue rising until the vapour pressure of the solvent and solute are the same. This increase in the pressure is the so-called "osmotic pressure of the solution."

§ 7. The Relation between the Boiling Point of a Solution and the Molecular Weight of the Solute.

In Figs. 62 and 97 the curve *PO* represents the vapour pressure of the solid, and *OQ* the vapour pressure of the pure liquid. The two curves

¹ Unless the pressure on one of the liquids exceeds 100 atmospheres.

intersect at the freezing point O . Let Q , Fig. 97, represent the boiling point of the solvent at 760 mm. pressure, then since the vapour pressure of a solution is less than the vapour pressure of the pure solvent, let $O'Q'$ represent the vapour pressure curve of a given solution. Then PM will represent the freezing point of the solvent, and PM' the freezing point of the solution. Since PM' is less than PM , the freezing point of the solution will be less than the freezing point of the solvent; and since PN' represents the boiling point of the solution and PN the boiling point of the solvent, the boiling point of the solution must be greater than the boiling point of the pure solvent. This agrees with experiment. For instance, with solutions of potassium iodide in 100 grams of water, G. T. Gerlach (1887) found:

Boiling point	100°	101°	102°	103°	104°	105°
Potassium iodide	0	15	30	45	60	74 grams

If the solutions are very strong the relation is not quite the same, but with dilute solutions, the raising of the boiling point of a dilute solution is directly proportional to the weight of the dissolved substance in a given weight of solvent. Double the concentration of the solution, and

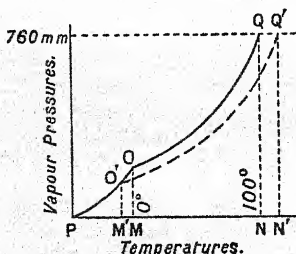


FIG. 97.

the elevation of the boiling point will be doubled. An equal number of molecules of the dissolved substance in the same quantity of a solvent give the same elevation of the boiling point—F. M. Raoult's law (1883-84). Hence the rise in the boiling point of a solvent is proportional to the number of molecules of the dissolved substance in solution, and inversely proportional to the molecular weight of the solute.

One gram-molecule of cane sugar (342 grams) dissolved in 100 grams of water raises the boiling point of the water 5.2° , that is, from 100° to 105.2° . This constant is called the boiling constant for water; it is sometimes called the "molecular elevation of the boiling point per 100 grams of solvent." Each solvent has its own specific boiling constant: e.g. acetone, 16.7; benzene, 26.7; ether, 21.6; carbon disulphide, 23.5, etc. The boiling constant is determined by finding the boiling point of, say, water and of aqueous solutions containing 0.02, 0.06, 0.10 gram-molecules of cane sugar, and calculating the results per 342 grams of cane sugar.

Suppose that w grams of a substance dissolved in 100 grams of water raised the boiling point of the water b° . Then, if M be used to denote the molecular weight of the substance, we have the proportion: $w : M = b : 5.2$; or, for substances dissolved in water:

$$\text{Molecular weight} = 5.2 \frac{w}{b}$$

This enables the molecular weight of many substances to be determined from their effect on the boiling point of water. The particular solvent to be used depends on the solubility of the substance under investigation. If ether is used in place of water, 5.2 must be altered to 21.6, etc.

EXAMPLES.—(1) E. Beckmann (1890) found that 2.0579 grams of iodine dissolved in 30.14 grams of ether raised the boiling point of the ether 0.566° . What is the molecular weight of iodine? Here, $2.0579 \text{ grams of iodine in } 30.14 \text{ grams of ether correspond with } 100 \times 2.0579 \div 30.14 = w = 6.8278 \text{ grams of iodine in } 100 \text{ grams of the solvent. Hence, } M = 21.6 \times 6.8278 \div 0.566 = 254.6$. This corresponds with the formula I_2 when iodine has a molecular weight of 253.84. The numbers seldom, if ever, coincide, but there can be no mistake in the significance of the figures.

(2) E. Beckmann (1890) found that a solution of 1.4475 grams of phosphorus in 54.65 grams of carbon disulphide raised the boiling point 0.486° . What is the molecular weight of the phosphorus? Answer: Molecular weight, 129.16. The atomic weight of phosphorus is 31, hence the molecule of phosphorus is represented: P_4 .

(3) A. Helft (1893) found that 0.2096 gram of sulphur in 17.79 grams of carbon disulphide raised the boiling point 0.107° . Hence show that the molecular weight of sulphur is probably S_8 . Here $w = 1.17$; and the molecular weight is 259. This is close to the theoretical value 256 for S_8 .

E. Beckmann's process for the determination of boiling points (1888-96).—The apparatus consists of a glass boiling tube *A*, Fig. 98, with a piece of platinum wire sealed in the bottom, and packed with beads to prevent irregular boiling. A side tube with a condenser *C* liquefies the vapour given off during the boiling; and the exposed end of the condenser is closed with a calcium chloride tube *D*. The boiling tube is surrounded by a jacket of some non-conducting material, *E*, to prevent the radiation of heat. The boiling tube is fitted with a Beckmann's thermometer, *T*, which can be read to $\frac{1}{100}$ of a degree, and set¹ so that the mercury is about half-way up the stem when the solvent is boiling. The boiling tube has a stoppered side tube, *J*, for introducing the solution under investigation. The whole is clamped to a stand and rests on an asbestos tray *F*.

The boiling point of the solvent is first determined. The boiling tube is weighed. The solvent is introduced and its boiling point determined

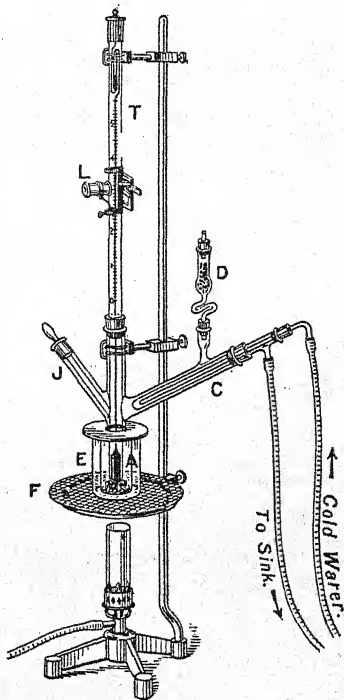


FIG. 98.—Beckmann's Apparatus for Boiling-point Determinations.

¹ This thermometer has a reservoir of mercury at the top so that it can be set for use at any desired temperature as indicated in text-books of laboratory processes. In this way, an inconveniently long, or an inconveniently large number of thermometers are not needed. The thermometer is always tapped before a reading to make sure the mercury is not lagging behind. The lens *L* facilitates the reading of the thermometer.

when the boiling is brisk and vigorous.¹ A known weight of the substance is then introduced, and the boiling point of the solution determined. A correction is made by subtracting 0.2 to 0.4 gram from the weight of the solvent in order to allow for the solvent condensed on to the walls of the apparatus and the condenser. The actual correction depends upon the nature of the solvent and the particular form of the apparatus used. The difficulty with this apparatus is to avoid fluctuations of temperature in the boiling tube due to the radiation of heat; dripping of the cold liquid from the condenser into the boiling solution, etc. Many other forms of apparatus for this determination have been devised.

Landsberger's method for the determination of boiling points (1898).—When the solution of non-volatile solute boils, the vapour of the solvent and solute are in equilibrium; this condition can be established by leading the vapour of the boiling solvent into the solution. When the solution is at its boiling point, the vapour will pass through the system without

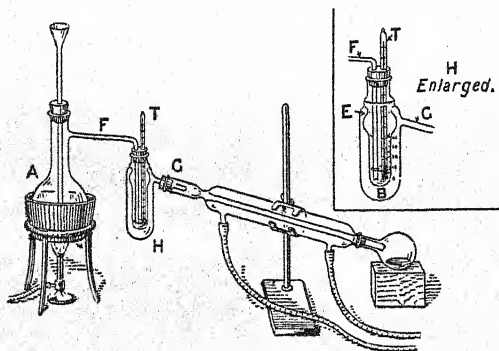


FIG. 99.—Landsberger's Apparatus for Boiling-point Determinations.

condensation, and if the solution is below this temperature some vapour will condense, and the latent heat of condensation will continue heating the solution until the boiling point is reached. There is virtually no danger of superheating the solution. This principle has been employed successfully in measuring the boiling points of solutions in molecular weight determinations. In W. Landsberger's apparatus, a modification of which is shown in Fig. 99, the solvent is boiled in the flask A, and the vapour passed into the solution *viâ* the tube F. The temperature of the solution is raised to its boiling point by the latent heat of condensation of the vapour of the solvent. The vapour of the solvent passes to the condenser G through E around the boiling tube, and thus the inner tube is jacketed with the vapour of the boiling solvent. This reduces radiation losses. The boiling point of the solvent is first determined, and a weighed amount of the solute is introduced into the inner tube B, which is graduated so that the boiling can be interrupted for a moment before more solute is added, and the volume of the solution read at a glance. With the preceding notation, with water as a solvent,

$$\text{Molecular weight} = 5.2 \frac{w}{b}$$

where w denotes the weight of the substance per 100 c.c. of the solvent,

¹ The barometer should be read to make sure no appreciable change occurs during a determination.

and b represents the elevation of the boiling point. If other solvents be used 5.2 is altered thus: for acetone, 22.2; benzene, 32.8; ether, 30.3; carbon disulphide, 26. If the boiling tube be weighed so that the amount of solvent is determined by weight, and not by volume, the original formula, on p. 256, is used.

EXAMPLE.—If 0.829 gram of a substance with 8.1 c.c. of acetone gave a rise of 1.47° in the boiling point of the solvent, what is the molecular weight of the substance? Here, $w = 100 \times 0.829 \div 8.1 = 10.24$. In the above formula, 5.2 for water is changed to 22.2 for acetone. Hence the desired molecular weight is $22.2 \times 10.24 \div 1.47 = 154$.

In the laboratory, advantage is taken of the fact that the boiling point of a solution is higher than the boiling point of the pure solvent to get liquids for baths, etc., boiling a few degrees higher than water, by dissolving the necessary amount of a salt in water. For example, a saturated solution of sodium nitrate boils at 120°, and a saturated solution of sodium chloride at 180°.

§ 8. The Relation between the Freezing Point of a Solution and the Molecular Weight of the Solute.

Similar remarks apply *mutatis mutandis* to the freezing point of solutions as were made with reference to the boiling point. A study of Fig. 97 will show that if the vapour pressure of solution is less than that of the pure solvent, the vapour pressure curve will cut the ice curve at a temperature below the freezing point of the pure solvent. This means that the freezing point of a given solution will be lower than the freezing point of the pure solvent, and experiment shows that the lowering of the freezing point will be proportional to the weight of the substance dissolved in a given weight of the solvent. This reminds us of Blagden's law, p. 192. The depression in the freezing point is proportional to the weight of the dissolved substance in a given weight of the solvent; and inversely proportional to the molecular weight of the dissolved substance.

Equal gram-molecules of different substances in the same solvent depress the freezing point to the same extent—F. M. Raoult's law (1883-84). A solution of sugar (342 grams), methyl alcohol (32.03 grams), etc., in 100 grams of water depresses the freezing point 18.5°. This is the freezing constant for water. It is also called "the molecular depression of the freezing point per 100 grams of solvent." Each solvent has its own specific freezing constant: e.g. acetic acid, 38.88°; benzene, 49°; mercury, 425°; naphthalene, 69°, etc.

If w grams of a substance, molecular weight M , dissolved in 100 grams of solvent, lower the freezing point f° , we have the proportion $w : M = f : 18.5$ for water; or, for substances dissolved in water,

$$\text{Molecular weight} = 18.5 \frac{w}{f}$$

This enables the molecular weight of a substance to be computed from its effect on the freezing point of water. The particular solvent to be selected is of course determined by the solubility of the substance under

investigation, and the number 18.5 must be replaced by another if a different solvent be used.

EXAMPLES.—(1) W. Tammann (1889) found that a solution of 0.022 gram of sodium in 100 grams of mercury lowered the freezing point of mercury 0.39° . What is the molecular weight of sodium? Here, $M = 425 \times 0.022 \div 0.39 = 23.8$. Hence the atomic and molecular weights are the same.

(2) W. R. Orndorff and J. White (1893) found that a solution of 0.2735 gram of hydrogen peroxide in 19.86 grams of water lowered the freezing point of water 0.746° . What is the molecular weight of hydrogen peroxide? Here $w = 100 \times 0.2735 \div 19.86 = 1.3773$; $f = 0.746$; hence, $M = 34.2$. This corresponds with the molecule H_2O_2 . An earlier determination by W. Tammann (1889) gave H_4O_4 , but this was afterwards found to be due to the use of an impure sample.

(3) J. Hertz (1890) found that 2.423 grams of sulphur in 100 grams of naphthalene lowered the freezing point of naphthalene 0.641° ; hence show that the molecular weight of sulphur under these conditions corresponds with the formula: S_8 . Answer: The molecular weight by experiment is 262, and by calculation for S_8 , 256.

E. Beckmann's process for the determination of freezing points.—Freezing-point determinations are usually made in Beckmann's apparatus.

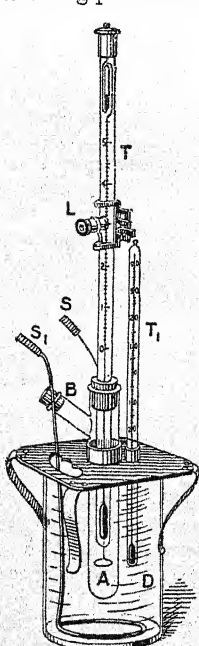


FIG. 100.—Beckmann's Apparatus for Freezing-point Determinations.

The tube *A*, Fig. 100, with a side neck, *B*, is weighed, and about 15 c.c. of the solvent are added, and the tube is weighed again. The Beckmann's thermometer, reading to the $\frac{1}{100}$ of a degree, and set so that the mercury is near the top of the scale when set for the freezing point of the solvent, has a reading lens. The thermometer *T* and a stirrer *S* are placed in the solvent, and the whole arrangement is placed in a glass tube *A* which serves as an air jacket. This is surrounded by a vessel *D* of water or some liquid at a temperature about 5° below the freezing point of the solvent. This vessel is fitted with a thermometer *T*₁ and stirrer *S*₁. The temperature recorded by the thermometer slowly falls until the solvent begins to freeze; it usually falls from 0.2° to 0.3° below the freezing point of the solvent, and then begins to rise to the freezing point proper. The thermometer should always be tapped before a reading is taken to make sure the mercury is not lagging behind. The highest point reached by the mercury in the thermometer is taken to be the freezing point of the solvent. Owing to undercooling, it is sometimes difficult to start the freezing of the solution. In that case, a few pieces of platinum foil, or a minute fragment of the frozen solvent, will start the freezing. It is sometimes necessary to introduce a correction for undercooling as indicated in text-books for the laboratory. Each determination should be repeated two or three times and the successive observations should agree within 0.002° to 0.003° . When the freezing point of the solvent has been determined, add a sufficient amount of the substance under investigation to give a depression of 0.3° to 0.5° .

After the freezing point has been determined again, find the freezing point after adding a second and then a third portion of the substance under investigation.

The molecular weights of volatile substances relative to the weights of the hydrogen molecule have been determined from the vapour density determinations and Avogadro's hypothesis. The osmotic pressure and related properties of solution enable the molecular weights of liquids and solids in solution to be determined. There is an extensive choice of solvents and it is possible to utilize such widely different solvents as stearic acid, mercury, ether, fused metals, etc. Molten salts containing water of crystallization have been used—*e.g.* sodium sulphate, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$; calcium chloride, $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$; lithium nitrate, LiNO_3 ; sodium chromate, $\text{Na}_2\text{CrO}_4 \cdot 10\text{H}_2\text{O}$; etc.—and the results are similar to those obtained with water as a solvent.

The molecular weights of a great many substances in solution are in agreement with those furnished by the vapour density method, yet there are some irregularities. The molecular weights of substances in solutions are sometimes greater and sometimes less than what we should expect. The results are then said to be abnormal. Ravault first assumed that organic substances must form double molecules in solution while the values for inorganic salts are normal. He then tried if extreme dilution would break down the molecules assumed to be doubled, and found it did not. Hence some other hypothesis must be sought in explanation of the phenomena.

§ 9. Anomalous or Abnormal Results for the Molecular Weights of Substances in Solution.

When a fact appears to be opposed to a whole train of deductions, it invariably proves to be capable of bearing some other interpretation.—SHERLOCK HOLMES.

We never profit more than by those unexpected results of experiments which contradict our analogies and theories.—L. B. GUYTON DE MORVEAU.

The meaning of "abnormal" in science.—We sometimes say that a phenomenon "ought to take place," but it does not. We have just used the word "abnormal" and "anomalous." These terms are not very happily chosen, and, as indicated on p. 102, they are sometimes used rather carelessly. The terms are not intended to imply that nature is erratic, arbitrary, and lawless. The words simply mean that in groping for the truth, an unexpected result has been obtained, which once stood, or now stands, challenging investigators to show how the unexpected should have been expected. In this sense it can be said that abnormal phenomena do not occur in nature. Some of the most treasured generalizations in science have been won by investigating the "abnormal." This applies in the laboratory as well as in the study. "I thank God," said H. Davy, "that I was not made a dexterous manipulator, for the most important of my discoveries have been suggested to me by my failures."

Revision of the gas equation $pV = RT$.—Let us return to the gas equation,

$$\frac{pV}{T} = \frac{p_1V_1}{T_1}$$

Remembering that the density D of a gas is equal to the molecular weight M divided by the volume v , or, $M = Dv$, we get $p/TD = p_1/T_1D_1$ when $M = M_1$. Let M , D , and v respectively denote the molecular weight, density and volume of the gas in one condition of temperature and pressure; and M_1 , D_1 , and v_1 , the same constants for another condition of temperature and pressure, we obtain, by substitution in the preceding equation: $pv/MT = p_1v_1/M_1T_1$. If we take the volume v_1 at some standard temperature T_1 and pressure p_1 , the numbers p_1 , v_1 , and T_1 will always have one fixed value. Let R denote this constant value of p_1v_1/T_1 . The gas equation then assumes the form:

$$pv = \frac{M}{M_1} RT; \text{ or, } pv = nRT$$

where n stands in place of the ratio of the molecular weights of the gas in the two conditions, M/M_1 . If the molecules of the gas neither dissociate nor polymerize when the conditions change, $M = M_1$, or $pv = RT$, because $n = 1$. But if the gas molecules polymerize or condense so that, say, two molecules combine together to form one molecule, there will be only half as many molecules in a given space as before; $M = \frac{1}{2}M_1$, and $pv = \frac{1}{2}RT$. If, however, the gas dissociates or decomposes so that each molecule of the gas forms two molecules of another gas or gases, then $M = 2M_1$, and we have $pv = 2RT$. Hence the ordinary gas equation: $pv = RT$, is a special case of the more general relation: $pv = nRT$, where the numerical value of n indicates whether or not the gas keeps the same molecular concentration during the change. If $n = 1$, there is neither dissociation nor polymerization; if n be less than unity, the gas polymerizes; and if n be greater than unity, the gas dissociates when the conditions are changed.

If the molecules of a dissolved substance are the same as the molecules would be if the substance were in the gaseous condition, the relation between the pressure, temperature, and concentration will be represented by the expression, $pv = nRT$; or, since the concentration c is inversely as the volume, by:

$$\frac{p}{c} = nRT$$

As before, if n be unity, the molecules of the substance in solution and in the gaseous condition are presumably similar; if n be greater than unity, the molecules dissociate when they pass into solution; and if n be less than unity, the molecules polymerize. If we apply the uncorrected relation, $p/c = RT$, it is now easy to see that if n be greater than unity (dissociation), the osmotic pressure will appear too high; and if n be less than unity (polymerization), the osmotic pressure will appear too low. When we speak of the lowering of the osmotic pressure, we also imply that the vapour pressure is increased, the boiling point is lowered, and the freezing point raised; and conversely, the raising of the osmotic pressure implies that the boiling point is raised, and the vapour pressure and freezing point are lowered.

Abnormally low osmotic pressures—polymerization of solute.—The depression of the freezing point of a solution of alcohol in benzene is just about half what we should expect if the molecules of alcohol were represented by the regular formula: C_2H_5OH . This means that the molecules

of alcohol— C_2H_5OH —in benzene solutions are doubled, and the molecule of alcohol is accordingly $C_4H_{10}(OH)_2$ in benzene solution. This phenomenon is common with molecules possessing hydroxyl, *i.e.* OH groups. Formic— $H.CO.OH$ —and acetic— $CH_3.CO.OH$ —acids, and indeed water, behave in a similar manner, and we know that if these acids be vaporized, they appear to have twice the molecular weight of what would obtain if their molecules could be really represented by the ordinary formulæ. Otherwise expressed, the molecules are polymerized. It is also necessary to remember that if the dissolved substance freezes out along with the solvent so as to form a kind of solid solution, the freezing point of the solution will be lower than that calculated from the regular molecular formula of the dissolved substance. Sometimes, indeed, the freezing point actually rises. There are many examples—solutions of lead, cadmium, tin, and gold in mercury; antimony in tin; etc.

Abnormally high osmotic pressures—dissociation of solute.—A very considerable number of aqueous solutions of acids, bases, and salts furnish a much greater osmotic pressure than we should naturally expect. The deviation of a gas from Avogadro's law is usually explained by assuming that the molecules of the gas are dissociated into simpler forms. Iodine molecules, I_2 , at high temperatures appear to behave as symbolized: $I_2 = I + I$. S. Arrhenius (1887) sought to explain the deviations of the molecular weights of salts, acids, and bases in aqueous solutions by assuming that the molecules are dissociated into simpler parts. The molecules of sodium chloride, for instance, are supposed to be dissociated in aqueous solutions into two parts—Na and Cl. The idea came as a surprise, and much opposition has been raised against this interpretation of the results, because there are no signs of chemical action which might be expected if the molecule of sodium chloride were dissociated into Na and Cl on solution in water. Accordingly, other hypotheses have been invented to make the first hypothesis fit the facts. In spite of this, Arrhenius' hypothesis at once explains in a seductive and plausible manner the abnormally high osmotic pressures obtained for these substances. There is a strange coincidence. Arrhenius determined the value of n —the number of molecules in the above equations for ninety different substances. He noticed at once that these substances could be roughly divided into two classes: those which gave values of n nearly unity were either non-conductors or poor conductors of electricity; whereas those which gave values of n materially greater than unity were fair or good conductors of electricity. In the following table n may be taken to represent, within the limits of experimental error, the relative number of molecules formed when one molecule of the substance is dissolved.

TABLE IX.—NORMAL AND ABNORMAL OSMOTIC PRESSURES.

Non-conductors.		Conductors.	
Substances in solution.	n	Substances in solution.	n
Methyl alcohol	0.94	Calcium nitrate	2.48
Mannite	0.97	Magnesium sulphate	1.25
Cane sugar	1.00	Strontium chloride	2.69
Ethyl acetate	0.96	Potassium chloride	1.81
Acetamide	0.96	Lithium chloride	1.92

We naturally inquire: What connection, if any, subsists between the alleged dissociation of the molecules of a substance in a solution and the conduction of electricity? How can one molecule of sodium chloride, one molecule of lithium chloride, and of hydrogen chloride, each furnish what appears to be two molecules when dissolved in water?

Questions.

1. An aqueous solution of LiCl, containing 8.5 grams in 1000 c.c., boils at 101.97° C. (760 mm.). What is the percentage dissociation of the LiCl? (Atomic weights: Li = 7, Cl = 35.5. Molecular raising of boiling point of water = 5.2° C.)—*Worcester Polytechnic Inst., U.S.A.*

2. 7.20 grams of a substance dissolved in 100 grams of water gave rise to an osmotic pressure of 9.65 atmospheres at 22°. Calculate the molecular weight of the substance.—*Sheffield Univ.*

3. A solution of 6.3 grams of a non-electrolyte in water, total volume 1000 c.c., freezes at -0.279° C. What is the molecular weight of the substance in solution? (Molecular lowering of the freezing point of water = 18.6.)—*Worcester Polytechnic Inst., U.S.A.*

4. What is meant by the term "osmosis"? Describe exactly what happens when a vessel with a semipermeable wall containing a solution of such a substance as sugar is placed in pure water. What is (a) the effect of increasing the strength of the sugar solution; and (b) the effect of raising the temperature of the whole apparatus?—*Univ. North Wales.*

5. What do you understand by the expressions "additive," "constitutive,"¹ and "colligative"? Illustrate your answer by examples.—*St. Andrews Univ.*

6. How have the freezing points of dilute solutions been exactly observed? What is the bearing of these observations on chemical theory?—*New Zealand Univ.*

7. The freezing point of an aqueous solution is -1.5° C. Find the relative lowering of the vapour pressure of the solution. (The molecular depression for water is 18.70.)—*St. Andrews Univ.*

8. Describe some experiment showing the phenomenon of "osmotic pressure." By what means has a relation been observed between the molecular weight of certain soluble substances and the osmotic pressure which they are supposed to exert? If the lower part of a U-tube be filled with a solution of sugar, and then pure water be gently poured into one limb so that the upper surface in one limb is pure water, and in the other limb is a solution of sugar, is osmotic pressure exerted, and if not, why not?—*New Zealand Univ.*

9. What do you understand by the term "osmotic pressure"? Describe in outline any two processes, one direct, the other indirect, for measuring osmotic pressure. Explain carefully how the molecular weight of a substance in solution can be determined when the osmotic pressure which it sets up is known.—*Board of Educ.*

10. What is osmotic pressure? Has this any connection with the pressure of a gas? Alcohol is said to be normal in regard to its vapour pressure and its osmotic pressure; ammonium chloride has an abnormal vapour pressure and osmotic pressure. Explain the meaning of the terms "normal" and "abnormal" used here.—*Sydney Univ.*

11. Explain how it is that vegetables will cook faster when boiled in a concentrated solution of salt than when boiled in water alone.

12. Describe the effects produced by soluble and insoluble substances on the boiling and melting points of water. What explanation can be given of the differences observed when common salt and sugar are respectively dissolved in sufficient pure water to form dilute solutions of the same molecular concentration? To what other properties of solutions does this explanation apply?—*Punjab Univ.*

13. Explain the principle of Raoult's method of determining molecular weights by observation of the freezing point. Give some account of its applications, and the chief results obtained.—*London Univ.*

¹ CONSTITUTIVE PROPERTIES have not yet been discussed. In these the preponderating factor is the mode of grouping of the atoms within the molecule. Examples will be indicated later—see "Isomerism."

CHAPTER XIII

CHLORINE AND HYDROGEN CHLORIDE.

§ 1. The Stassfurt Salt Beds.

The agricultural world is almost wholly dependent upon a very limited region in North Germany for its supply of potash used as an essential ingredient in certain fertilizers.—W. C. BLASDALE.

THE remarkable deposits of potassium, magnesium and sodium salts in the country around Stassfurt, in Prussian Saxony, may be very roughly divided into four strata, illustrated diagrammatically in Fig. 101.

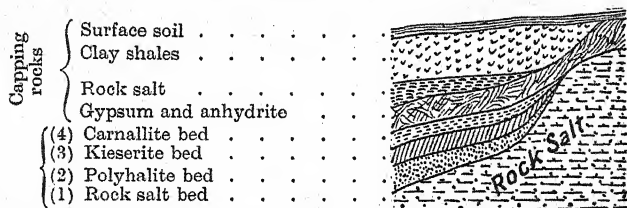


FIG. 101.—Diagrammatic Geological Section of a Part of the Stassfurt Salt Bed.

(1) *Rock salt bed*.—An immense basal bed of rock salt, broken up at fairly regular intervals with 2 to 5 inch bands of *anhydrite*— CaSO_4 .

(2) *Polyhalite bed*.—Above the basal salt is a layer of rock salt, sometimes 200 feet thick, mixed with bands of magnesium chloride and *polyhalite*— $2\text{CaSO}_4 \cdot \text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$.

(3) *Kieserite bed*.—Resting on the polyhalite bed is a layer of rock salt, sometimes 100 feet thick, mixed with layers of *kieserite*— $\text{MgSO}_4 \cdot \text{H}_2\text{O}$ —and other sulphates, about 1 foot thickness.

(4) *Carnallite bed*.—Finally comes a reddish layer of rock salt associated with masses of *kainite*— $\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$; *carnallite*— $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$; and a few other salts of magnesium and potassium, e.g. *sylvine*— KCl and *leonite*— $\text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$.

These deposits are capped by layers of gypsum— $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ —and anhydrite— CaSO_4 ; rock salt; bunter clay shales; and finally the surface soil. In addition to gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$), anhydrite (CaSO_4), and rock salt (NaCl), the principal salts found in the Stassfurt deposits are:

Sylvine	KCl
Carnallite	$\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$
Kieserite	$\text{MgSO}_4 \cdot \text{H}_2\text{O}$
Schönite	$\text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$
Kainite	$\text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$
Polyhalite	$\text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 2\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$

With several other salts of lesser importance: *e.g.*, *astrakanite* ($\text{Na}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 4\text{H}_2\text{O}$); *boracite* ($2\text{Mg}_3\text{B}_3\text{O}_{15} \cdot \text{MgCl}_2$); *glauberite* ($\text{Na}_2\text{SO}_4 \cdot \text{CaSO}_4$); *leonite* ($\text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$); etc.

History.—In the thirteenth century, salt springs—called “sool”—were well known in the region of Stassfurt, and a certain amount of salt—sodium chloride—was obtained from them, but the springs were abandoned when rock salt was discovered in other parts of Germany. About 1840, borings were made in the hope of finding deposits which would give a good quality of salt, and about 1850, shafts were sunk, and rock salt mined. The material excavated with the rock salt was rejected

as worthless, and called *Abraumsalze*—German: *Abraum*, refuse; *Salze*, salts. As a result of the investigations of Rose and Rammelsberg, the *Abraumsalze* were recognized to be a valuable source of potassium and magnesium. Processes were then devised for the extraction of the potassium and magnesium salts, with the result that the rock salt became of little value, and the *Abraumsalze* became of primary importance. A. Frank erected the first works for the extraction of potassium chloride in 1861, and an important industry, controlled by the “German Kali Syndicate,” has been established. The Stassfurt salt deposits have been the subject of elaborate investigations by J. H. van’t Hoff and his pupils in the light of the phase rule of J. W. Gibbs.

Crystallization of salts from mixed solutions.—The simple cases of the crystallization of a solution saturated with but one salt was discussed on p. 36; and of solutions of two salts which do not react with one another, nor form hydrates, was discussed on p. 203. The phenomenon is more complex when the salts present in the solution form a series of hydrates, or when the salts

can react with one another to form double salts. A solution of potassium chloride and magnesium chloride not only furnishes crystals of magnesium chloride, $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$, and of potassium chloride, KCl ; but also crystals of the double salt, $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ —*carnallite*. Referring to Fig. 102, the line *AB* represents the effect of additions of potassium chloride on the amount of magnesium chloride required to form a saturated solution at 25° . The solubilities are here expressed in terms of gram-molecules of MgCl_2 and of K_2Cl_2 per 1000 c.c.¹ The line *BC*

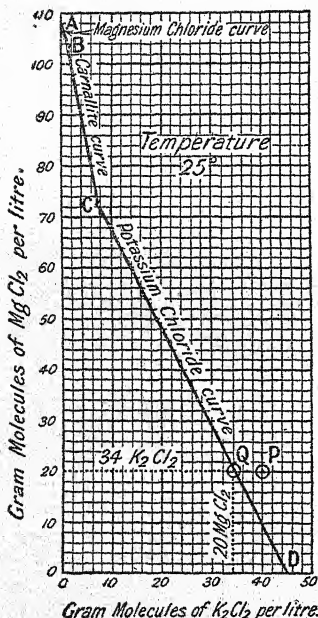


FIG. 102.—Solubilities of Magnesium and Potassium Chlorides in Mixed Solutions.

¹ “ K_2Cl_2 ” is written in place of “ 2KCl ” without any implication that the molecule of potassium chloride is K_2Cl_2 . This is done to keep equivalent molecules of magnesium and potassium chlorides as units for ordinates and abscissae.

represents the amount of carnallite in a saturated solution as the amount of potassium chloride is increased; and the line CD represents the amount of potassium chloride in a saturated solution as the amount of magnesium chloride is increased.

All solutions represented by points on the lines $ABCD$ are saturated with one of these salts. If therefore we start with a solution containing, say, 20 gram-molecules of magnesium chloride, and 40 gram-molecules of K_2Cl_2 per litre, the composition of the solution will be represented by a point P on the diagram. Crystals of potassium chloride will be deposited until the composition of the solution is represented by a point Q on the line CD , that is, potassium chloride will be deposited until the solution contains 20 gram-molecules of magnesium chloride, and 34 of K_2Cl_2 per litre. If the solution be concentrated by evaporation at 25° , potassium chloride will continue separating until the composition of the solution can be represented by a point C , that is, until the solution contains about $5\frac{1}{2}$ gram-molecules of K_2Cl_2 ; and 72.5 gram-molecules of $MgCl_2$ per litre. If the mother liquid be still further concentrated at 25° , crystals of carnallite and of potassium chloride will separate until the concentration of the solution is represented by a point B , corresponding with one gram-molecule of K_2Cl_2 , and 105 gram-molecules of $MgCl_2$. Any further concentration of the mother liquid will lead to the separation of magnesium chloride and carnallite in constant proportions until the solution is dry.

If the temperature at which the crystals are removed be different, different results will be obtained, because of differences in the solubilities of the different salts at varying temperatures; the formation of hydrates at temperatures above or below transition points, etc. The same principles obtain even with still more complicated examples, say a mixture of potassium chloride and magnesium sulphate, where we have the reaction, $2KCl + MgSO_4 = MgCl_2 + K_2SO_4$. This solution may lead to the separation of crystals of potassium chloride and sulphate; magnesium chloride; two hydrates (" $6H_2O$ " and " $7H_2O$ ") of magnesium sulphate; carnallite, $KCl.MgCl_2.6H_2O$; and schönite— $K_2SO_4.MgSO_4.6H_2O$.

Origin.—It is generally thought that the Stassfurt beds are of marine origin, and have been formed by the natural evaporation of water, during countless years, in an inland prehistoric sea, probably communicating with the ocean by a shallow bar. The sea must have been intermittently replenished by water bringing in more salts, as could occur when driven over the bar by high tides and gales; there must also have been a number of geological elevations and depressions to account for the succession of strata. The order in which salts are deposited from the evaporation of sea-water is very nearly the same as the geological succession observed at Stassfurt. Neglecting the calcium sulphate, the evaporation of sea-water furnishes successively: (1) a deposit of sodium chloride; (2) sodium chloride mixed with magnesium sulphate; (3) sodium chloride and leonite; (4) sodium chloride, leonite, and potassium chloride, or sodium chloride and kainite; (5) sodium chloride, kieserite, and carnallite; (6) sodium chloride, kieserite, carnallite, magnesium chloride; and (7) the solution dries without further change.

Uses.—The Stassfurt salts furnish magnesium salts which are used for the preparation of magnesium and its salts. The potash salts are largely used as manures in agriculture; and the potassium chloride is used as a

basis for the manufacture of the many different kinds of potassium salts used in commerce—carbonate, hydroxide, nitrate, chloride, chromates, alums, ferrocyanide, cyanide, iodide, bromide, etc. Chlorine and bromine are obtained from the mother liquids by electrolytic and other processes. Boric acid and borax are prepared from boracite. Cæsium and rubidium are recovered from crude carnallite and sylvine.

§ 2. Sodium and Potassium Chlorides.

Thou shalt offer salt with all thy meat offerings.—LEVITICUS.

Salt is frequently mentioned in the Bible. The religious rites of the Jews and of many other nations testify to the high esteem and sanctity in which the ancients held salt. The modern Arabians are said to still practise the covenant of salt, but only on occasions of the highest importance. The superstitious custom of throwing a little spilt salt over the shoulder to propitiate the fairies is a survival of the old superstition when salt was included among the offerings to the gods. Sodium chloride, commonly called "salt," is considered to be an essential constituent of animal food. It is said that there are few things more distressing than "salt hunger"; and in the dreaded "salt torture" of the Chinese, the prisoner is provided with ample food but insufficient salt. One writer estimates that about 29 lbs. of salt per head of population per annum is used directly or indirectly with the food for man. The 0.1 per cent. of hydrochloric acid present in the gastric and mucous fluids of the alimentary canal is derived from the decomposition of the salt taken in with the food. Plant-eating animals get much of the salt they require from grass and leaves; herbivorous animals have been known to travel hundreds of miles to a "salt-lick" or "salt spring" in order to satisfy their craving for salt. Carnivorous animals get their salt from the blood of the animals on which they feed.

The occurrence of salt.—Rock salt occurs in transparent or translucent cubic crystals, either colourless or varying in tint from white, to dirty grey, to yellow, to reddish yellow; and sometimes blue or purple. It is sometimes called *halite*. Rock salt is found in Nantwich, Northwich, Middlewich (Cheshire); Droitwich¹ (Worcestershire); Stassfurt (Prussian Saxony); Cardona, Castile (Spain); California, Utah, Kansas, New York, Virginia, Ohio, Michigan (United States); and numerous other places. The mines at Wieliczka (Galicia, Austria) have been worked continuously for 600 years. The salt deposit is said to be 500 miles long, 20 miles broad, and 1200 feet thick. The galleries and chambers in this mine extend over 30 miles in length and yield 55,000 tons per annum. The salt town in the mine includes a market-place, river, and church. A comparatively large amount of salt is dissolved in sea-water, and in the water of many salt springs, and salt wells. The water of the Mediterranean Sea, for instance, contains 3.37 per cent. of solids in solution. The composition of sea salt from different parts of the world approximates

NaCl	KCl	MgCl ₂	CaSO ₄	MgSO ₄	(Mg,Ca)CO ₃	(Na ₂ ,Mg)Br ₂
73.58	3.71	9.61	4.66	5.42	0.1	1.0%

¹ The names of these localities indicate the antiquity of the salt industry, since, in Saxon times, a place where salt was dug was called a "wich."

It has been estimated that the rivers of the world discharge about 160 million tons of salt into the sea every year, and that the seas contain 144 billion tons of salt in solution, "enough indeed to cover the whole of the dry land to a depth of 400 ft." Consequently, under present conditions, it must have taken at least 90 million years to accumulate the amount of salt now present in the seas.

The separation of salt from sea-water.—Countries not supplied with rock salt must either import salt from more favoured countries, or resort to the concentration of sea-water, or the water of salt springs. Evaporation is not an expensive process in warm countries, or where coal is cheap; for instance, in the *salterns* or salt gardens on the shores of the Mediterranean Sea the sea-water is concentrated by evaporation in large shallow tanks—"salterns"—exposed to wind and sun. As the solution—"brine"—becomes concentrated, the crystals of salt—*bay salt*—which separate are lifted out by means of perforated shovels, and allowed to drain beside the evaporation tanks. The crystals are allowed to stand in heaps exposed to occasional showers, whereby much of the magnesium chloride is leached out. The product is sometimes called *solar salt*. The mother liquid—"bittern"—was once used for the manufacture of bromine. In cold countries, *e.g.* on the shores of the White Sea (Russia), the sea-water is concentrated by freezing (p. 192). Ice first separates; and the residual brine is further concentrated by evaporation over a fire.

The purification of rock salt.—Rock salt is often mined by bringing the solid salt to the surface in lumps. In some salt beds, the salt is mined by forcing water into the beds *via* a well made for the purpose. The brine is afterwards pumped to the surface, and the liquid concentrated in salterns; or by allowing the solution to trickle from elevated tanks over ricks of brushwood—"graduator's"—so arranged that the solution is fully exposed to the prevailing winds. The liquid may be afterwards concentrated by evaporation in shallow pans heated artificially, particularly if fuel is cheap. As the salt crystallizes out, it is removed by means of perforated shovels. If much calcium sulphate be present, it will separate first. It must therefore be removed before the salt. The potassium and magnesium salts separate last. Sometimes a little milk of lime is added to precipitate the calcium and magnesium salts. In *multiple evaporators*, the brine is evaporated in closed vessels so that the steam from one vessel heats the brine in the adjacent one. The first one of the series is heated by a series of pipes through which steam passes, and the last of the series is connected with an exhaust pump whereby the brine is made to boil at a reduced temperature. This last is a vacuum pan. The salt from the vacuum pans is fine-grained; for coarse-grained salt, open air evaporation is needed. If fine-grained "table salt" is needed, the brine is evaporated rapidly near its boiling point; but for the manufacture of coarse-grained "fine-salt" the evaporation is conducted slowly at a comparatively low temperature (45°) so as to get the salt in comparatively large crystals. The salt obtained by the evaporation of sea-water will be contaminated with small quantities of other salts as impurities: calcium chloride, magnesium chloride, calcium sulphate, and magnesium sulphate. Cheshire salt, for instance, contains about 98.3 per cent. of sodium chloride; the remaining impurities are mainly insoluble matter, calcium sulphate, and magnesium and calcium chlorides. The presence of magnesium chloride

makes salt very hygroscopic, so that it becomes lumpy and moist; some "damp-proof" salts are mixed with, say, 3 per cent. of bone-ash to get over the difficulty to some extent.

The extraction of potassium chloride from carnallite.—Potassium chloride occurs as sylvine—KCl—in the Stassfurt deposits, and associated with magnesium chloride in carnallite. Potassium chloride was formerly obtained from sea-water, but much of the potassium chloride of commerce is now made from carnallite— $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ —of the Stassfurt deposits. The crude carnallite contains about 61 per cent. of carnallite, 25 per cent. of kieserite, 12 per cent. of rock salt, and 2 per cent. of anhydrite and clay. This salt is crushed and digested in large tanks with the mother liquid left from preceding operations. This liquid contains chiefly magnesium chloride. The mixture is heated by blowing steam into the liquid. The potassium chloride readily dissolves in this liquid, while most of the sodium chloride and magnesium sulphate, associated with the crude carnallite, remain as an insoluble residue. The liquid is allowed to settle for an hour, and then decanted into large iron vats, where crystals containing 64 to 69 per cent. of potassium chloride are deposited. The impurities are mainly sodium chloride (20–22 per cent.), magnesium chloride (7.5–8.5 per cent.), and 0.4 per cent. of magnesium bromide and calcium sulphate. The principle underlying the process will appear from the study of Fig. 102. The crystals of potassium chloride so obtained are washed in cold water so as to remove the more soluble sodium chloride. This process yields a product containing 84 to 98 per cent. of potassium chloride, according to the number of washings. The further purification of the salt involves a re-solution and re-crystallization. The mother liquors are washed up for more salts, and finally used for the extraction of a fresh lot of crude carnallite.

The purification of sodium and potassium chlorides.—Sodium as well as potassium chloride can be purified by adding concentrated hydrochloric acid to a cold concentrated aqueous solution of the respective salts; better results are obtained by passing gaseous hydrogen chloride through the salt solutions. The impurities remain in solution while the chlorides are precipitated in a very fair state of purity.

Properties.—Both chlorides crystallize in cubes, and the two salts are isomorphous. The crystals are anhydrous. A little water may be mechanically entangled with the crystals, which causes the salts to decrepitate when heated. Sodium chloride melts at 801° , and potassium chloride is said to melt between 759° and 766° . Both salts sublime at higher temperatures without decomposition. Sodium chloride boils at about 1750° . If sodium chloride is heated in a tube filled inside with a porous cell containing air, the inside of the cell soon fills with chlorine gas, while the tube containing the sodium chloride vapour contains very little free chlorine. The reaction between the clay and the sodium chloride vapour decomposes the latter, liberating chlorine gas. The solubilities of the two salts in water were discussed on pp. 189 and 204. A comparison of the related chlorides of lithium, sodium, potassium, rubidium, and caesium, shows that they all crystallize in cubes when anhydrous—Fig. 66, left. Their solubilities, expressed in grams per 100 c.c. of water at 15° , are:

Lithium.	Sodium.	Potassium.	Rubidium.	Caesium.
80	36	33.4	80	very high

and the solubilities in alcohol follow the same rule, sodium and potassium chlorides being almost insoluble in alcohol.

Composition.—The composition of both salts has been established in the same manner. By analysis, J. S. Stas found sodium chloride contained 39.39 per cent. of sodium, and 60.61 per cent. of chlorine. Hence, after division by the respective atomic weights of these elements, we get the atomic ratio $\text{Na} : \text{Cl} = 1 : 1$ corresponding with the formula $(\text{NaCl})_n$. The difficult volatility of sodium chloride—contrasted with, say, mercuric chloride—suggests a complex molecule for the solid, but the X-radiograms of the crystals do not support this contention. W. Nernst (1903) found the vapour densities of both sodium and potassium chlorides at 2000° corresponded respectively with the formulæ NaCl and KCl .

Uses of sodium chloride.—Salt is used for seasoning food—table salt. Salt for table use should be free from magnesium and calcium chlorides, for these substances make salt very deliquescent—particularly in moist weather. Salt is also used for preserving meat, fish, etc.; in the manufacture of sodium salts, soaps, etc.; in glazing common pottery—drain pipes, some sanitary goods, etc.—“salt glaze”; and also in the manufacture of chlorine compounds, directly or indirectly. The alkali industries start with salt as the raw material.

§ 3. Hydrogen Chloride—Preparation and Properties.

Molecular weight, $\text{HCl} = 36.47$; melting point, -112.5° ; boiling point, between -83° and -84° ; critical temperature, 52.3° . Vapour density ($\text{H}_2 = 2$), 36.49; (air = 1) 1.269. One litre weighs 1.641 grams under normal conditions. Specific gravity of liquid, 0.908 at 0° .

Preparation.—When sodium or potassium chloride is treated with warm dilute sulphuric acid (1 : 1) in a flask (Fig. 103), a gas is given off. The gas is very soluble in water and it cannot be collected over water, but it can be collected over mercury. In general laboratory work, it is often convenient to collect relatively heavy gases by the upward displacement of air. The gas was once called the “spirit of salt,” but is now called “hydrogen chloride” and symbolized “HCl.” The reaction, at 18° with equimolar proportions of salt and acid of sp. gr. 1.84, is represented: $\text{H}_2\text{SO}_4 + \text{NaCl} = \text{NaHSO}_4 + \text{HCl}$; at 120° a further amount of hydrogen chloride is given off: $\text{NaHSO}_4 + \text{H}_2\text{SO}_4 + \text{NaCl} = 2\text{NaHSO}_4 + \text{HCl}$. No heat change occurs when the salt and acid are mixed at 18° ; with potassium chloride the temperature rises to 30° , and with ammonium chloride it falls to about 1° . The gas can be dried by passage through wash-bottles containing concentrated sulphuric acid. If the concentrated acid be employed with an excess of sodium chloride, at a rather more elevated temperature, the reaction is represented by: $\text{H}_2\text{SO}_4 + 2\text{NaCl} = \text{Na}_2\text{SO}_4 + 2\text{HCl}$. The same gas is often made in the labora-

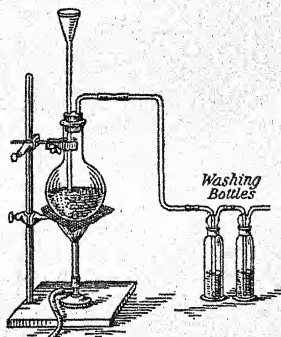


FIG. 103.—Preparation of Hydrogen Chloride.

tory at ordinary temperatures by mixing concentrated hydrochloric acid with an excess of sodium or ammonium chloride and dropping concentrated sulphuric acid from a tap funnel into the mixture, as indicated in Fig. 53.

Properties.—Hydrogen chloride is a colourless gas which irritates the mucous membrane when inhaled. It forms dense fumes in moist air. The gas is incombustible, and a non-supporter of combustion. Hydrogen chloride is extremely soluble in water: 1 c.c. of water at 0° dissolves 525 c.c. of the gas; and at 20°, 440 c.c. The heat of solution is $\text{HCl} + \text{Aq} = \text{HClaq} + 17.4 \text{ Cals.}$ The aqueous solution is strongly acid, and is called "hydrochloric acid," "spirits of salt," or "muriatic acid"—from the Latin, *muria*, brine. Hydrogen chloride partially dissociates into free chlorine and hydrogen at about 1500°: $2\text{HCl} \rightleftharpoons \text{H}_2 + \text{Cl}_2$. Hydrogen chloride is easily condensed to a colourless liquid by pressure. At 10° a pressure of 40 atmospheres will liquefy the gas, and at -16° , 20 atmospheres suffice. The liquid boils at -83.7° under ordinary atmospheric pressures. The liquid freezes to a white crystalline mass melting at -112.5° . The liquid does not act on many metals which are vigorously attacked by the aqueous solution of hydrogen chloride. Neither the dry liquid nor the dry gas acts on blue litmus. The gas is fairly stable at high temperatures, being dissociated only to the extent of 0.274 per cent. at 1537° .

The freezing temperatures of aqueous solutions of hydrogen chloride.—The freezing temperatures of aqueous solutions of hydrogen

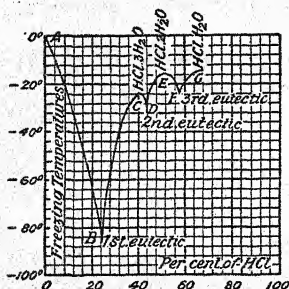


FIG. 104.—Freezing-point Curves of Mixtures of Hydrogen Chloride and Water.

chloride of different concentrations have been determined for solutions containing less than 67 per cent. of hydrogen chloride, as shown in Fig. 104—F. F. Rupert (1909). Starting with pure water, the addition of hydrogen chloride steadily depresses the freezing point to the eutectic temperature— 85° , *AB*, when the solution contains 25 per cent. of HCl. Further additions of hydrogen chloride raise the freezing temperature, *BC*, up to -24.4° , when the mixture contains 40.3 per cent. of HCl, and thus corresponds with the trihydrate— $\text{HCl.3H}_2\text{O}$. Continued additions of hydrogen chloride depress the freezing point curve, *CD*, to a second eutectic— 28° ,

and then raise it, *DE*, to a second maximum, -17.7° , corresponding with 50.31 per cent. of HCl, that is, with the dihydrate— $\text{HCl.2H}_2\text{O}$. The freezing-point curve again descends, *EF*, to a third eutectic, -23.5° , with increasing concentration, and rises, *FG*, to a third maximum, -15.35° , when the solution contains 66.9 per cent. of HCl, corresponding with the monohydrate— $\text{HCl.H}_2\text{O}$. With more concentrated solutions, the liquid separates into two layers on cooling. The first eutectic is concerned with the system $\text{H}_2\text{O} : \text{HCl.3H}_2\text{O}$; the second eutectic with the two hydrates $\text{HCl.3H}_2\text{O} : \text{HCl.2H}_2\text{O}$; and the third eutectic with the system $\text{HCl.2H}_2\text{O} : \text{HCl.H}_2\text{O}$. Each of these three systems behaves like ice and brine indicated in Fig. 58. The three maxima thus correspond with the three hydrates

$\text{HCl} \cdot 3\text{H}_2\text{O}$; $\text{HCl} \cdot 2\text{H}_2\text{O}$; $\text{HCl} \cdot \text{H}_2\text{O}$. All three hydrates have been isolated in the form of white crystalline solids. The existence of an octohydrate— $\text{HCl} \cdot 8\text{H}_2\text{O}$ has been inferred from the heat of solution of hydrogen chloride in water; but it is quite an imaginary hydrate, for it has not been isolated. If it does exist, its presence is not indicated on the freezing-point curve, Fig. 104. The maxima in a freezing-point curve thus represent compounds, and the minima, eutectic mixtures. The freezing or melting point curve of a mixture of two (or more) substances is often sensitive enough to demonstrate the existence of the more stable compounds, but it is too rough and inaccurate for the less stable compounds. Nevertheless, this method of investigating stable hydrates has been applied to ammonia, ferric chloride hydrates, perchloric acid, sulphuric acid, nitric acid, etc., and it has important applications in metallurgy.

The effect of hydrogen chloride on the vapour pressure of water.—The effect of hydrogen chloride on the boiling point of water is illustrated by the curve, Fig. 105.

H. E. Roscoe and W. Dittmar found that if an aqueous solution of hydrogen chloride containing more than 20.24 per cent. HCl be heated, hydrogen chloride with but little water is given off; the solution becomes less concentrated; the vapour pressure of the solution diminishes; and consequently, the boiling point rises as indicated by the curve. This continues until the solution contains nearly 20.24 per cent. of HCl, when its boiling point attains the maximum,

110°; any further boiling does not affect the concentration of the aqueous solution because dilute acid containing 20.24 per cent. of HCl distils unchanged. Again, if an acid containing less than 20.24 per cent. of HCl be boiled, water accompanied by a little hydrogen chloride passes off; the boiling point of the solution gradually rises; and the solution at the same time becomes more concentrated until it contains 20.24 per cent. HCl, when the acid distils over unchanged at 110°. Hence 110° is the maximum boiling point of hydrochloric acid at atmospheric pressures. Similar phenomena occur with nitric acid and with several other acids. It was once thought that the acid which corresponded with the maximum boiling point was an octohydrate—that is, a chemical compound of hydrogen chloride and water— $\text{HCl} \cdot 8\text{H}_2\text{O}$; but since the composition of the constant boiling acid varies with the pressure,¹ and since compounds

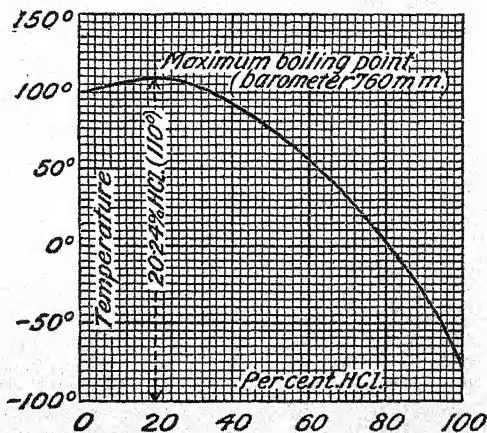


FIG. 105.—Boiling Points of Aqueous Solutions of Hydrogen Chloride.

¹ For instance, at 100 mm. pressure, the maximum boiling point is nearly 62°, and the constant boiling acid contains 22.8 per cent. HCl.

do not usually vary in composition with changes of pressure, this hypothesis has been abandoned. Still, under constant conditions, the composition of the constant boiling acid is always the same, and this fact has suggested a means of preparing acids of definite concentration for volumetric analysis. With solutions of oxygen, ammonia, hydrogen, and nitrogen in water, the more volatile constituent leaves the solution before all the boiling water has evaporated.

Fuming liquids.—Since concentrated aqueous solutions of hydrogen chloride have a vapour pressure greater than water, we can see a reason for the fuming of hydrochloric acid in air. We know, of course, that hot water "fumes" in air because the cooler air in the vicinity of the hot water is quickly saturated with water vapour. Water at ordinary temperatures does not fume because it cannot give off more vapour than the air at the same temperature can retain. Concentrated hydrochloric acid fumes because the vapours which are given off unite with the aqueous vapour in the atmosphere to form an acid in which the partial pressure of the water vapour is less than that of pure water at the same temperature. Consequently, the air in the vicinity of the concentrated acid is very quickly saturated with respect to the vapour of the new acid which is formed. The new acid, in consequence, condenses to minute globules of liquid which appear as mist. Dilute acids do not fume because any vapours which they give off do not form a liquid with a smaller partial pressure of water vapour than water itself. Hence, only those substances fume which give off vapours which unite with water to form a mixture or a compound with a smaller vapour pressure than water.

The formation of chlorides.—The aqueous solution of hydrogen chloride dissolves many metals, forming chlorides and liberating hydrogen, *e.g.* $\text{Zn} + 2\text{HCl} = \text{ZnCl}_2 + \text{H}_2$. Zinc, magnesium, iron, aluminium, and tin are readily dissolved by cold dilute acid—the action with aluminium and tin is, however, rather slow in the cold, but much quicker in hot concentrated acid. Mercury, silver, gold, and platinum are not dissolved by the hot or cold acid; copper and lead are not dissolved by the cold dilute acid unless exposed to the air, when the action is very slow, these metals are only slowly attacked by the hot concentrated acid. Hydrochloric acid reacts with oxides, hydroxides, and carbonates producing the corresponding chlorides. Most of the chlorides are easily dissolved by water. In qualitative analysis, it is usual to divide the metals into two groups: those with soluble and those with insoluble chlorides. The "insoluble" chlorides are: silver, mercurous, cuprous, aurous, thallous, and lead chlorides. Lead chloride is, however, appreciably soluble in cold water, and much more soluble in hot water. It therefore occupies a position midway between the soluble and "insoluble" chlorides. The chloride is often more readily volatile than many of the other compounds of a given metal.

Manufacture of hydrochloric acid.—Hydrochloric acid is obtained as a by-product in the manufacture of sodium carbonate from sodium chloride. In the first stage of the process, sodium chloride is treated with sulphuric acid, and the gas which is evolved is passed up stone towers filled with lumps of coke down which a stream of water trickles. The water absorbs the gas and is collected in suitable receivers at the base of the tower. Commercial hydrochloric acid may be contaminated with ferric chloride,

free chlorine, sulphurous and sulphuric acids, arsenic chloride, etc. The first-named impurity gives commercial hydrochloric acid its yellow colour.

Uses.—Hydrochloric acid is used in the manufacture of chlorine, in dyeing, calico printing, the manufacture of colours, phosphates, and as a general laboratory reagent. A carboy of the acid (sp. gr. 1.16) holds about 112 lbs.—the commercial acid sells at about 8s. per cwt., and the "pure acid" at about $2\frac{1}{2}d.$ per lb., pre-war prices. Crude commercial hydrochloric acid has about 30 per cent. of HCl by weight; and the "commercially pure" acid has a sp. gr. of about 1.2, and contains about 25–35 per cent. of HCl.

History.—Judging from the writings attributed to Geber, hydrochloric acid was known to the early Arabian chemists; but the preparation of the pure acid—*spiritus salis*—seems to have been first described by Basil Valentine (1644). The acid appears to have been made by distilling a mixture of common salt and green vitriol (ferrous sulphate). J. R. Glauber (1648) described the preparation of the acid by the action of sulphuric acid on rock salt. Stephen Hales (1727) noticed that a gas very soluble in water was made by heating sulphuric acid with sal ammoniac (ammonium chloride), and J. Priestley, about 1772, collected the gas over mercury. Priestley called the gas *marine-acid air* in reference to its mode of formation from sea-salt.

§ 4. The Action of Oxidizing Agents on Hydrogen Chloride.

The action of oxidizing agents on hydrogen chloride or hydrochloric acid is very interesting. For instance, K. W. Scheele (1774) found that when hydrochloric acid is heated with manganese dioxide, a yellowish-green gas, soluble in water, is given off. Scheele considered the yellowish-green gas to be muriatic acid freed from hydrogen (phlogiston), that is, in the language of his time, "dephlogisticated muriatic acid." A. L. Lavoisier (1789) named the gas oxymuriatic acid, or oxygenated muriatic acid, because he considered it to be an oxide of muriatic (i.e. hydrochloric) acid; and, consistent with his oxygen theory of acids, p. 167, Lavoisier considered muriatic acid to be a compound of oxygen with an hypothetical muriatic base—murium—this was later symbolized Mu, while hydrochloric acid was symbolized MuO_2 , and Scheele's acid MuO_3 . Hence, added Lavoisier, muriatic and oxymuriatic acids are related to each other like sulphurous and sulphuric acids. This certainly seemed to be a most plausible explanation of the reactions. Lavoisier's hypothesis was supported by an observation of C. L. Berthollet (1785), who noticed that an aqueous solution of Scheele's gas, the so-called oxymuriatic acid, when exposed to sunlight, gives off bubbles of oxygen gas, and forms muriatic acid.

J. L. Gay-Lussac and J. Thénard (1809) tried to deoxidize oxymuriatic acid, so as to isolate the hypothetical "muriatic base" of Lavoisier, by passing the dry gas over red hot carbon, but when the carbon was freed from hydrogen, the attempt to separate from oxymuriatic acid anything but itself was a failure. While favouring Lavoisier's hypothesis, Gay-Lussac and Thénard added: "the facts can also be explained on the hypothesis that oxymuriatic acid is an elementary body." Here, then, are two

rival hypotheses as to the nature of oxymuriatic acid—the yellowish-green gas discovered by Scheele!

In 1810 H. Davy tried, without success, to decompose oxymuriatic acid. He found that when hydrogen chloride is heated with metallic sodium or potassium, the metallic chloride, and hydrogen are formed, but neither water nor oxygen is obtained. Davy claimed that Scheele's view is an expression of the facts, while Lavoisier's theory, though "beautiful and satisfactory," is based upon a dubious hypothesis. The definition of an element (p. 25) will not permit us to assume that oxymuriatic acid is a compound, because, in spite of repeated efforts, nothing simpler than itself has ever been obtained from the gas. In order to avoid the hypothesis implied in the term "oxymuriatic acid," H. Davy proposed the alternative term "chlorine" and symbol "Cl"—from the Greek *χλωρος* (chloros), green. The term "chlorine" is thus "founded upon one of the obvious and characteristic properties of the gas—its colour." J. J. Berzelius, in 1811, thought Lavoisier was right, but in 1821 he gave up this view. Lavoisier's hypothesis died a lingering death, and Davy's view is now generally adopted. According to Davy's theory, Berthollet's observation is explained by the equation: $2\text{H}_2\text{O} + 2\text{Cl}_2 = 4\text{HCl} + \text{O}_2$; that is, the oxygen comes from the water, not from the chlorine. Similarly, the formation of chlorine by the action of oxidizing agents upon hydrochloric acid is due to the removal of hydrogen from hydrogen chloride. In symbols: $4\text{HCl} + \text{O}_2 = 2\text{H}_2\text{O} + 2\text{Cl}_2$.

§ 5. Chlorine—Preparation.

Atomic weight, Cl = 35.46; molecular weight, Cl_2 = 70.92. Generally univalent, occasionally quinque- and septivalent. Melting point, -102° ; boiling point -33.8° ; critical temperature, $+146^\circ$. Relative vapour density (H_2 = 2) 71.63; (air = 1) 2.49. One litre weighs 3.22 grams under normal conditions. Specific gravity of liquid at 0° , 1.47.

Chlorine gas is usually obtained, as indicated above, by the action of oxidizing agents—manganese dioxide, lead dioxide, barium dioxide, potassium dichromate, potassium permanganate, etc.—upon hydrochloric acid. Scheele, the discoverer of chlorine, used a mixture of manganese dioxide and hydrochloric acid; a mixture of sodium chloride, sulphuric acid, and manganese dioxide may also be used. The mixture is heated in the apparatus illustrated in Fig. 103. The action in both cases depends upon the formation of manganese trichloride, and the simultaneous oxidation of the hydrogen of hydrogen chloride: $2\text{MnO}_2 + 8\text{HCl} \rightarrow 4\text{H}_2\text{O} + 2\text{MnCl}_3 + \text{Cl}_2$. When the mixture is warmed, the manganese trichloride is decomposed: $2\text{MnCl}_3 \rightarrow 2\text{MnCl}_2 + \text{Cl}_2$. When this process is used on a manufacturing scale, the manganese chloride— MnCl_2 —is treated by W. Weldon's recovery process (1867), which is a modification of an earlier process by C. Dunlop (1855). Air is blown through the hot residual liquid which has been previously treated with an excess of "milk of lime." In this way a considerable amount of calcium manganite— $\text{CaO} \cdot \text{MnO}_2$ —is formed. Heat is liberated, and the action once started takes place without the application of any heat. According to F. Fischer, some manganous hydroxide, $\text{Mn}(\text{OH})_2$, and manganese oxydihydroxide, $\text{MnO}(\text{OH})_2$, are formed. The precipitate is called "Weldon's mud" or

"manganese mud." The presence of a base is necessary for the oxidation of the oxide of manganese by air. The mixture is allowed to settle, and the clear liquid run off. The "mud" is gradually run into the chlorine still containing hydrochloric acid so that the "manganese" is used over and over again.

Potassium permanganate is a very convenient oxidizing agent for preparing chlorine. A flask containing some crystals of potassium permanganate is fitted as indicated in Fig. 50, and connected with a wash-bottle containing concentrated sulphuric acid. Dilute hydrochloric acid is run, drop by drop, from a tap funnel, when chlorine is evolved by the reaction: $2\text{KMnO}_4 + 16\text{HCl} \rightarrow 8\text{H}_2\text{O} + 2\text{KCl} + 2\text{MnCl}_2 + 5\text{Cl}_2$. Chlorine is also made by the action of an excess of hydrochloric acid upon an alkaline hypochlorite or bleaching powder. The bleaching powder may be purchased compressed into cubes, with or without plaster of Paris, and used in Kipp's apparatus, Fig. 13, with hydrochloric acid (L. Winkler, 1887). The gas attacks mercury, and it is not, therefore, collected in the mercury trough; it is also fairly soluble in water and the solution is rather irritating. For general laboratory work the gas can be collected over hot water saturated with salt, or better, by the upward displacement of air in the stink closet.

Gold and platinum chlorides give off chlorine when heated, but these compounds are too expensive for the preparation of chlorine, except for very special work, such as V. Meyer's work on the vapour density of chlorine, where platinous chloride was used as the source of chlorine. Cupric chloride— CuCl_2 —also gives off chlorine when heated: $2\text{CuCl}_2 = 2\text{CuCl} + \text{Cl}_2$, and this method is sometimes used for the preparation of pure chlorine. In W. Weldon and A. R. Péchiney's process (1885) for chlorine, magnesium chloride is heated in a current of air. Magnesium oxide and free chlorine are produced: $2\text{MgCl}_2 + \text{O}_2 = 2\text{MgO} + 2\text{Cl}_2 - 14$ Cals. The oxide of magnesium is then treated with hydrogen chloride to regenerate magnesium chloride: $\text{MgO} + 2\text{HCl} = \text{H}_2\text{O} + \text{MgCl}_2$. The chloride so formed is again heated in a current of air, so that the process of manufacture of chlorine is continuous.

Deacon's process for chlorine.—The oxidation of hydrogen in hydrogen chloride can be effected by atmospheric oxygen, by passing the mixed gases through a tube at a high temperature. The action takes place below 400° in the presence of pumice-stone saturated with cuprous chloride— CuCl . The result of the reaction is represented by the equation: $4\text{HCl} + \text{O}_2 + \text{CuCl} = 2\text{H}_2\text{O} + 2\text{Cl}_2 + \text{CuCl}$. The cuprous chloride remaining at the end of the reaction has the same composition as at the beginning. It is supposed that the first action results in the formation of a copper oxychloride: $4\text{CuCl} + \text{O}_2 = 2\text{Cu}_2\text{OCl}_2$; followed by: $\text{Cu}_2\text{OCl}_2 + 2\text{HCl} = 2\text{CuCl}_2 + \text{H}_2\text{O}$; and finally by: $2\text{CuCl}_2 = 2\text{CuCl} + \text{Cl}_2$. The chlorine is necessarily contaminated with undecomposed hydrogen chloride, atmospheric nitrogen, atmospheric oxygen, and steam. The steam and hydrogen chloride can be removed by washing, etc. The chlorine so prepared is used in the manufacture of bleaching powder, where the presence of the impurities does no particular harm. This is the principle of H. W. Deacon's process (1868). The reaction can be illustrated by the apparatus shown in Fig. 106. Air is forced from a gas holder through a hot solution of concentrated hydrochloric acid.

The mixture of air and hydrogen chloride so obtained is passed through a wash-bottle containing water, and then through a hot porcelain tube containing pumice-stone impregnated with a solution of cupric chloride and dried. The chlorine gas obtained at the exit can be collected in the usual manner. It is of course mixed with the excess of air, nitrogen, etc. The reaction never runs completely to an end, but approaches a state of equilibrium which fixes a definite limit to the yield of chlorine which can be obtained at any particular temperature and concentration of the reacting gases. In the reaction: $4\text{HCl} + \text{O}_2 = 2\text{Cl}_2 + 2\text{H}_2\text{O} + 26 \text{ Cals.}$, both chlorine and oxygen are competing for the hydrogen; at 577° both appear equally strong, for the hydrogen is distributed equally between the chlorine and oxygen. At higher temperatures the chlorine is stronger than oxygen, because less free chlorine is obtained than at lower temperatures, when the affinity of oxygen for the hydrogen is the stronger. In conse-

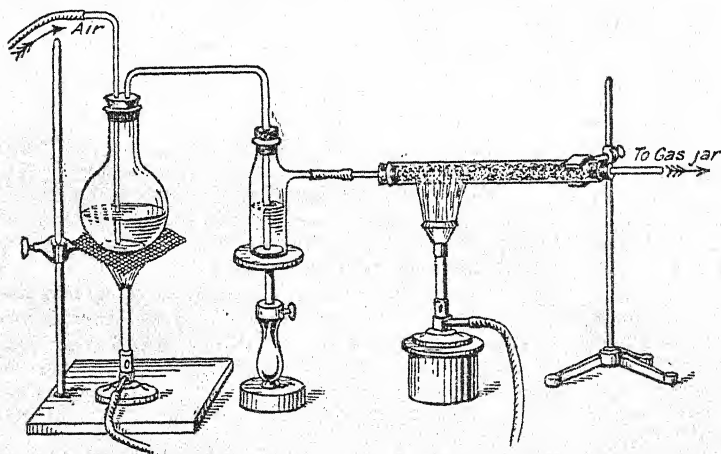


FIG. 106.—Illustration of Deacon's Process for Chlorine.

quence, a greater yield of free chlorine is obtained at temperatures lower than 577° . The temperature, however, cannot be reduced indefinitely because the reaction would then become inconveniently slow, even in the presence of the catalytic agent—cuprous chloride. The catalytic agent begins to volatilize at temperatures even below 430° .

Electrolytic processes for chlorine and alkaline hydroxides.—If an aqueous solution of potassium chloride or sodium chloride be electrolyzed, chlorine (anion) appears at the anode, and the metal (cation) at the cathode. In the case of sodium chloride, we have $\text{NaCl} = \text{Na} + \text{Cl}$. The metal then reacts with the water, liberating hydrogen and forming sodium hydroxide. The net result of the electrolysis is: $2\text{NaCl} + 2\text{H}_2\text{O} = \text{Cl}_2 + \text{H}_2 + 2\text{NaOH}$, so that hydrogen gas appears at the cathode and chlorine at the anode. The two electrodes must be separated to prevent the sodium hydroxide formed at the cathode mixing with the chlorine discharged at the anode. The separation is effected:—

(1) *Diaphragm process.*—By using a porous diaphragm—Portland

cement, earthenware, asbestos, limestone, etc. This permits electrolytic conduction, and prevents the solutions mixing but very slowly—P. Matthes and Weber, 1886.

(2) *Bell process*.—By enclosing the anode in an inverted non-conducting bell with the cathode outside—W. Bein, 1893.

(3) *Mercury cathode process*.—The sodium is dissolved by the mercury to form an amalgam. The amalgam is removed from the cell and treated with water, when sodium hydroxide and mercury are obtained. The mercury is returned to the cell to be used over and over again—E. Solvay's process, 1898.

(4) *Mercury diaphragm process*.—By the use of a mercury diaphragm as in H. Y. Castner's process, 1893, illustrated in Fig. 107. Castner's cell has three compartments. The two outer compartments are fitted with graphite anodes (+); and the middle compartment is fitted with an iron grid (−) to serve as cathode. The non-porous partitions do not reach quite to the bottom of the cell but dip into a layer of mercury covering the bottom. A solution of alkali chloride flows through the two outer cells, and water through the inner compartment. The brine in the outer compartment is decomposed by the electric current into chlorine at the anode and sodium at the cathode. The latter dissolves in the mercury, at the cathode, and the chlorine at the anodes escapes *via* the exit pipes. The sodium amalgam diffuses into the inner chamber, and there, coming into contact with the water, is immediately decomposed into sodium hydroxide and mercury. The hydrogen escapes through the loosely fitting cover. The sodium hydroxide is run into a special tank as required. A slow rocking motion is imparted to the cell during the electrolysis, by an eccentric wheel, so as to make the mercury flow from one compartment to the other along the bottom of the cell.

(5) *Fused electrolyte*.—In Acker's process (1898), now abandoned, molten lead was used as anode for collecting the sodium. The alloy of lead and sodium was decomposed by steam to form hydrogen and sodium hydroxide. Fused alkaline chlorides also furnish chlorine when electrolyzed. Pure chlorine, for special experiments, has been made by the electrolysis of pure fused silver chloride with carbon electrodes. For the electrolysis of hydrochloric acid, see p. 283.

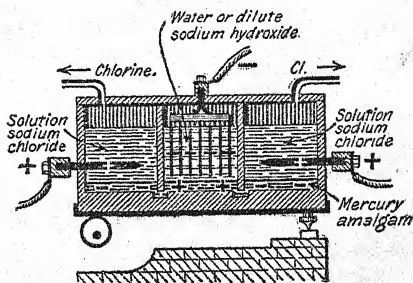


FIG. 107.—Castner's Mercury Diaphragm Process.

§ 6. Chlorine—Properties.

Chlorine is a yellowish-green gas with an irritating smell. It attacks the membrane of the throat, lungs, and nose. If a little be breathed for some time, it causes an irritating cough attended by the spitting of bloody mucus. Larger quantities of chlorine are fatal. Chlorine is about

$2\frac{1}{2}$ times as heavy as air, and being slightly soluble in cold water, it can be collected by the upward displacement of air, over hot water or over a concentrated solution of salt in which it is not so very soluble.

The action of chlorine on water.—100 volumes of water at 0° dissolve 461 volumes of the gas, and at 20° , 226 volumes. The solution in water—chlorine water—is yellowish-green in colour, and it has the taste and smell of chlorine itself. Chlorine water is conveniently made by passing the

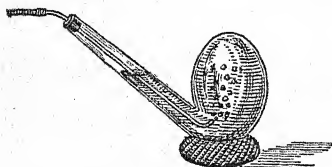


FIG. 108.—Preparation of Chlorine Water.

gas into an inverted retort, placed as illustrated in Fig. 108. When the water is cooled by surrounding it with melting ice, yellow rhombic octahedral crystals of chlorine octohydrate— $\text{Cl}_2 \cdot 8\text{H}_2\text{O}$ —separate. The hydrate decomposes slowly at temperatures just over 0° , and rapidly at higher temperatures. Thus the vapour pressure of the hydrate at 0° is

250 mm. ; at 5° , 481 mm. ; and at 10° , 832 mm. If the hydrate be sealed in a tube, it will melt at ordinary temperatures, forming two liquids—an upper aqueous layer, and a lower layer containing most of the chlorine.

The liquefaction of chlorine.—On March 5, 1823, M. Faraday was operating with chlorine hydrate in a sealed tube. Dr. J. A. Paris called at the laboratory and noticed some oily matter in the tube Faraday was using ; he rallied Faraday “upon the carelessness of employing soiled vessels.” Faraday started to open the tube by filing the sealed end ; the contents of the tube suddenly exploded, and the “oil” vanished. Faraday repeated the experiment, and Dr. Paris, next morning, received the laconic note :—

“Dear Sir,—The *oil* you noticed yesterday turned out to be liquid chlorine. —Yours faithfully, MICHAEL FARADAY.”

Chlorine can be condensed to a golden-yellow liquid at 0° and 6 atmospheres pressure. By sealing chlorine hydrate in one limb of a Λ -shaped

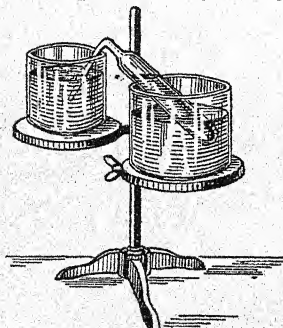


FIG. 109.—Liquefaction of Chlorine.

tube, and placing that leg in warm water while the other leg is immersed in a freezing mixture (Fig. 109) of, say, ice and salt, yellow oily drops of liquid chlorine condense in the cold limb. The liquid boils at -33.6° under atmospheric pressure, and freezes to a pale yellow crystalline mass which melts at -102° . Liquid chlorine is sold commercially in steel cylinders. Experiments with compressed gases are dangerous. Many serious and disastrous explosions are on record. When Thilorier was experimenting with liquid carbon dioxide, a cast iron vessel exploded and

“tore off both legs from the unfortunate M. Hery.”

The action of chlorine on other elements.—Thoroughly dry chlorine is somewhat inert chemically, and it has no appreciable action upon bright metallic sodium, copper, etc. Moist chlorine is particularly active.

Chlorine does not combine with oxygen directly although several compounds of chlorine and oxygen can be obtained indirectly. The inert gases, nitrogen, oxygen, carbon, and some of the rarer platinum metals, resist attack by *free* chlorine. The direct union of many of the elements with chlorine is attended by incandescence—for instance, powdered antimony, arsenic, and bismuth when shaken into a flask containing chlorine. Since the chlorides of antimony, etc., so formed are poisonous, the experiments are best made in a closed system, illustrated Fig. 110. When the bulb tube containing the powdered element is raised, it is easy to shake the contents through the flask of chlorine to illustrate incandescence which attends the combustion without an escape of the poisonous chlorides into the atmosphere of the room. Copper, brass foil, Dutch metal, phosphorus, boron, and silicon also ignite spontaneously in chlorine. Molten sodium, hot brass wire, and iron wire also burn in chlorine.

In sunlight, equal volumes of hydrogen and chlorine combine with explosion. The same remark applies when a mixture is exposed to the light of burning magnesium. Let a jar of hydrogen and a similar jar of chlorine be placed mouth to mouth (Fig. 36), the gases well mixed, the jars separated and immediately covered with greased glass plates. When a lighted taper is applied to the mouth of one jar, the gases unite with explosion. A piece of cold but recently ignited charcoal placed in the other jar frequently causes the gases to combine with explosive violence.

Hydrogen does not appear to combine with chlorine with appreciable velocity in the dark, but in diffused daylight, the two slowly combine to form hydrogen chloride. The speed of the reaction is proportional to the intensity of the light. Hence, actinometers have been designed to measure the intensity of light in terms of the speed of combination of a mixture of hydrogen and chlorine gases. If light be filtered through a layer of chlorine gas before it impinges on the mixture of hydrogen and chlorine, the light produces no appreciable effect. Insolated chlorine rises slightly in temperature, even when the heat rays have been filtered from the incident light. It is therefore inferred that *actinic (light) energy absorbed by chlorine is at once degraded into thermal energy. If hydrogen be associated with the chlorine, when exposed to light, the process of degradation of the actinic energy is accompanied by chemical action* (J. W. Mellor, 1902). This mode of describing the facts does not tell us very much. The presence of minute traces of impurities in the gases retards the rate of combination in a remarkable way.

A jet of burning hydrogen lowered into a jar of chlorine continues burning with the formation of hydrogen chloride. Chlorine gas may also be burnt in an atmosphere of hydrogen. Hydrocarbons are decomposed by chlorine; for instance, a piece of cotton wool soaked in warm turpentine ($C_{10}H_{16}$) will inflame when placed in a jar of chlorine. The burning of the turpentine is accompanied by the separation of dense clouds of free carbon. The chlorine combines with the hydrogen forming hydrogen



FIG. 110.—Combustion of Antimony, etc., in Chlorine.

chloride, and carbon is set free. A wax candle burns in chlorine with a very smoky flame; the hydrocarbon—wax—is decomposed in a similar manner. See also Fig. 317. Hence, chlorine may be regarded as non-combustible, and a supporter of combustion.

Oxidizing effects of moist chlorine.—Moist chlorine, or chlorine water, is a powerful oxidizing agent. We have seen that chlorine water is decomposed in sunlight, p. 276. Oxygen gas is given off and hydrogen chloride is formed: $2\text{H}_2\text{O} + 2\text{Cl}_2 = 4\text{HCl} + \text{O}_2$. If a piece of coloured litmus paper, coloured petals of a flower, or a piece of cloth dyed with turkey red or indigo blue be placed in a jar of dry chlorine no appreciable change occurs; but if moisture be present, the colours are bleached by the chlorine. The action appears to be due to the formation of a colourless oxidation product. Ordinary oxygen will not do the work of bleaching. One school of chemists therefore assumes that the oxidation is effected by the nascent oxygen—p. 345; another, that it is due to the decomposition of hypochlorous acid formed by the action of chlorine on water. The oxidation of sulphur dioxide by moist chlorine is of interest because it results in the simultaneous formation of both hydrochloric and sulphuric acids: $\text{Cl}_2 + \text{SO}_2 + 2\text{H}_2\text{O} = \text{H}_2\text{SO}_4 + 2\text{HCl}$.

Dissociation of the chlorine molecule.—The density of chlorine below 600° corresponds with 70.92, and the molecule is accordingly Cl_2 ; at higher temperatures, the density assumes lower values. At 1200° , the density is 47.3, that is, about two-thirds the normal value for chlorine. This is explained by the assumption that the two-atom chlorine molecules commence to dissociate into one-atom molecules above 600° : $\text{Cl}_2 \rightleftharpoons \text{Cl} + \text{Cl}$.

Uses.—Chlorine is used in the manufacture of bleaching powder; in the extraction of gold; in bleaching—e.g. wood pulp, etc.; in the manufacture of bromine, etc.

§ 7. The Composition of Hydrogen Chloride and the Atomic Weight of Chlorine.

1. **Analysis of hydrogen chloride by sodium amalgam.**—A stoppered glass tube—about 70 cm. long and 1.5 cm. in diameter—is filled with dry hydrogen chloride over mercury. Sodium amalgam is then dropped into the tube, and the tube *immediately* closed with its stopper. The tube is inverted several times in succession, and then opened while its mouth is dipping under mercury. Mercury rushes into the tube, and the residual gas is brought under atmospheric pressure by raising or lowering the tube until the mercury inside and outside is at the same level (Fig. 111). The volume of the residual gas is noted. The residual gas can then be tested in the usual manner. It is hydrogen. The hydrogen chloride reacts with the sodium of the mercury amalgam forming sodium chloride and liberating hydrogen. The object of using sodium amalgam in place of metallic sodium is one of convenience. This experiment demonstrates that hydrogen chloride contains half its own volume of hydrogen. Hence, from Avogadro's hypothesis, one molecule of hydrogen chloride contains half a molecule, that is, one atom of hydrogen. The formula is therefore HCl_x , where x represents the number of atoms of chlorine in the molecule. The vapour density of hydrogen chloride is nearly 36.5 ($\text{H}_2 = 2$). Hence the molecular weight is 36.5, and the weight

of hydrogen in the molecule is 1. The molecule of hydrogen chloride thus contains $36.5 - 1 = 35.5$ parts of chlorine. This is the atomic weight of chlorine. Hence the formula for hydrogen chloride is HCl .

2. The electrolysis of hydrochloric acid.—When concentrated hydrochloric acid is electrolyzed, a mixture of equal volumes of hydrogen and chlorine is obtained. Carbon electrodes are used because the chlorine slowly attacks platinum. Chlorine gas is also soluble in concentrated hydrochloric acid, so that the acid should be saturated with chlorine before the attempt is made to measure the gaseous products of electrolysis. If the vessel containing the acid is arranged so that the acid about one electrode is connected with the acid about the other electrode by a glass tube junction, Fig. 112, it is only necessary to saturate the solution about one electrode with chlorine. The apparatus illustrated, Fig. 112, devised by L. Meyer, has two limbs filled with concentrated hydrochloric acid, and an electric current is passed until the liquid in the chlorine limb is saturated with chlorine. The two gas receivers are put into communication with the electrolytic vessel by suitably turning the three-way stopcocks. The gas receivers have, of course, been previously filled with liquid—a saturated solution of sodium chloride—by placing a dish of the liquid below each

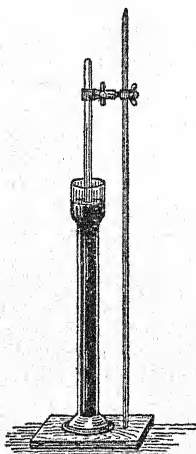


FIG. 111.—Volume Composition of Hydrogen Chloride.

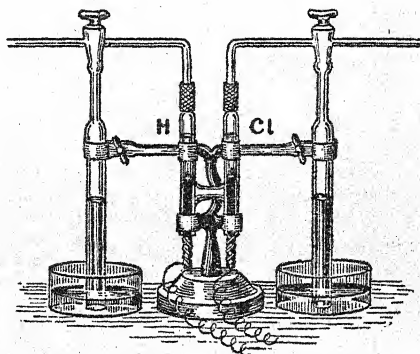


FIG. 112.—Meyer's Apparatus for the Electrolysis of Hydrochloric Acid.

receiver and applying suction at the proper exit tube when the three-way cocks are suitably turned. The gases collect in the tubes at equal rates. The experiment shows that during the electrolysis of concentrated hydrochloric acid, the volume of hydrogen liberated at the one electrode is equal to the volume of chlorine liberated at the other electrode. Assuming that the hydrogen chloride dissolved in the water is alone decomposed by the electric current, it follows that hydrogen chloride contains equal volumes of hydrogen and of chlorine, and therefore also an equal number of atoms; or the formula is H_xCl_x , where x is evaluated as before from the vapour density 36.5. This demonstration of the composition

of hydrogen chloride, though interesting as circumstantial evidence, is not a *proof* unless supported by accessory evidence. A similar demonstration applied to the analogous hydrofluoric acid would "prove" that hydrogen fluoride is a compound of hydrogen and oxygen.

When concentrated hydrochloric acid is diluted with eight volumes of water and electrolyzed, some oxygen is evolved along with the chlorine; with nine volumes of water, still more oxygen is evolved. The more dilute the acid the greater the amount of oxygen, until, with water acidified with a few drops of acid, no chlorine, but oxygen alone is obtained at the anode, p. 67.

3. The synthesis of hydrogen chloride.—The mixed gases obtained by the electrolysis of concentrated hydrochloric acid in an apparatus resembling Fig. 19 are passed through a stout glass "explosion" tube with a stopcock at each end, A, Fig. 113. The tube is packed with lime and glass wool to absorb the chlorine. Instead of this the exit tube may lead to the fume closet. When all the air is displaced, the stopcocks are closed. One of the stopcocks may be opened while the corresponding end of the tube is dipping under concentrated sulphuric acid; no gas enters or leaves the apparatus. The tube and contents are exposed to sunlight or to the light from burning magnesium. The face must be protected in case the tube should burst during the explosion. When the tube is cold, open one of the stopcocks while the corresponding end is dipping under concentrated sulphuric acid; no gas enters or leaves the tube. This shows that

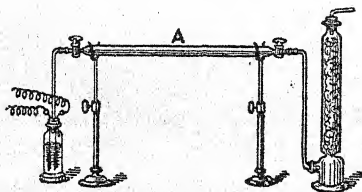


FIG. 113.—Collecting Electrolytic Hydrogen and Chlorine.

no change in volume has taken place as a result of the explosion. It can be proved that the tube contains nothing but hydrogen chloride by opening the tip of the tube under water. The hydrogen chloride will be absorbed and water will rise and fill the tube except for a little air (or perhaps a slight excess of hydrogen) which might have been present. This experiment shows that one volume of hydrogen unites with one volume of chlorine to form two volumes of hydrogen chloride.

Hydrogen chloride contains the equivalent of half its volume of chlorine and half its volume of hydrogen, or, by Avogadro's hypothesis, assuming the hydrogen and chlorine each contain two atoms, one molecule of hydrogen chloride contains half a molecule of hydrogen and half a molecule of chlorine, that is, one molecule of hydrogen chloride contains an atom of chlorine and an atom of hydrogen. The formula is therefore HCl. This agrees with the vapour density determination of hydrogen chloride which furnishes 36.49 ($H_2 = 2$). If the atomic weight of chlorine be 35.46, and of hydrogen 1.008 ($O = 16$), it follows that the formula for hydrogen chloride is HCl.

4. The atomic weight of chlorine.—The combining weight of chlorine has been deduced by finding how much silver chloride can be obtained from a given amount of silver. The results show that $Ag : Cl = 107.88 : 35.46$. The ratio $H : Cl$ has also been determined by the combustion of hydrogen in chlorine in an apparatus similar in principle to that

employed by Morley for the combustion of hydrogen in oxygen, p. 62. Collecting the best determinations, it is found that the ratio varies between 1 : 35·43 and 1 : 35·46. The best representative value is taken to be 35·46. Again collecting together the vapour density determinations of all known volatile chlorides, we obtain a table from which the following has been abstracted :

TABLE X.—MOLECULAR WEIGHTS OF VOLATILE COMPOUNDS.

Volatile chloride.	Vapour density.	Formula of compound : molecular weight = vapour density.	Amount of chlorine in the molecule.
Hydrogen chloride . . .	36·5	HCl	35·46
Chlorine	70·9	Cl ₂	70·92
Mercuric chloride . . .	273·6	HgCl ₂	70·92
Arsenic trichloride . . .	182·1	AsCl ₃	106·38
Tin tetrachloride . . .	260·2	SnCl ₄	141·84
Phosphorus pentachloride .	208·3	PCl ₅	177·30

The smallest combining weight of chlorine in any one of these compounds corresponds with the combining weight 35·46—oxygen = 16—and accordingly this number is taken to represent the atomic weight of chlorine. The atomic and equivalent weights of chlorine have the same numerical value.

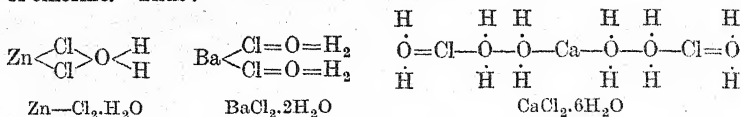
§ 8. The Chlorides of Calcium, Barium, Magnesium, and Zinc.

Many metals readily combine with chlorine, forming chlorides, and often in several different proportions : *e.g.* stannous chloride, SnCl₂, and stannic chloride, SnCl₄ ; cuprous chloride, CuCl, and cupric chloride, CuCl₂ ; mercurous chloride, HgCl, and mercuric chloride, HgCl₂ ; ferrous chloride, FeCl₂, and ferric chloride, FeCl₃, etc.

In addition to the methods of making chlorides previously described, namely, by acting on the metal, the metallic oxide, hydroxide, or carbonate with hydrochloric acid ; and by adding a soluble chloride to a solution of the metallic salt whereby a sparingly soluble chloride is precipitated, chlorides can be made by the action of chlorine on the metallic oxide, or a mixture of the metallic oxide with carbon, or by the action of chlorine or hydrogen chloride on the metal, and by some other methods to be described later. The last-named processes are generally used when an anhydrous chloride is wanted.

Calcium, strontium, barium, magnesium, zinc, cadmium, and beryllium chlorides form a closely graded series. They can all be made by the general methods, and they all crystallize from aqueous solutions united with water which varies from one to seven molecules of water per molecule of the chloride. For instance, ZnCl₂·H₂O ; BaCl₂·2H₂O ; CaCl₂·6H₂O, etc. The mode of combination of the water molecules, generally called water of crystallization, is not clear. It is sometimes supposed to depend upon

the quadrivalency of the oxygen of the water molecule, and the tervalency of chlorine. Thus:



If the heats of combination of the different elements with chlorine be plotted with the atomic weights of the elements, a curious periodic curve is obtained as indicated in Fig. 114, where the dotted lines denote the want of data. A similar periodic curve is obtained when the heats of formation of the oxides are treated in a similar way.

Double chlorides.—The chlorides may also unite with other chlorides forming double chlorides, *e.g.* $\text{MgCl}_2 \cdot \text{NH}_4\text{Cl} \cdot 6\text{H}_2\text{O}$; $\text{MgCl}_2 \cdot \text{KCl} \cdot 6\text{H}_2\text{O}$; $\text{ZnCl}_2 \cdot \text{NH}_4\text{Cl}$, etc., when their mixed aqueous solution are concentrated and allowed to crystallize. The double chlorides can be dried without

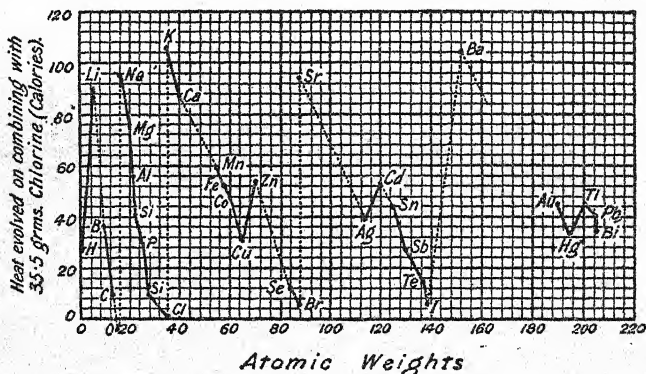
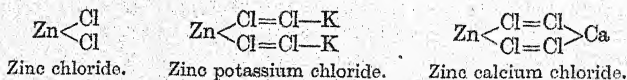


FIG. 114.—Heats of Formation of the Chlorides (A. P. Laurie, 1882).

decomposition. This property is not so well defined with the calcium, barium, and strontium compounds as with beryllium, zinc, cadmium, and magnesium compounds. The property has been attributed to the multivalency of chlorine. Thus, with tervalent chlorine:



Neutral salt reactions.—If a concentrated solution of neutral sodium sulphate be heated with red litmus, the blue coloration shows that an *alkaline solution* is formed, the alkalinity disappears as the solution cools. The reaction is not shown when a solution of sodium chloride is treated in a similar manner, but if nitroanthrochrysone sulphonic acid is used as indicator, sodium chloride too shows the reaction. If rosaniline just decolorized with dilute sodium hydroxide be used as indicator, both solutions give a magenta coloration indicating an *acid reaction*. With the so-

called basic indicators, the above-mentioned salts—as well as others—give alkaline reactions, and with acid indicators, an *acidic reaction*. Hence it is inferred that the hydrolysis of these neutral salts $\text{NaCl} + \text{H}_2\text{O} \rightleftharpoons \text{HCl} + \text{NaOH}$ is too feeble to be detected under ordinary conditions, but it becomes evident when an indicator is used which can itself unite with one product of the hydrolysis.

Hydrolysis.—These chlorides are slightly decomposed in aqueous solutions. The term “hydrolysis” is generally employed for the interaction between a salt and water whereby free acid and free base, or an acid and a basic salt, are formed. Hydrolysis is thus a kind of reversion of the process of neutralization of an acid with a base, or of base with an acid. In the case of magnesium chloride and water, a certain amount of the magnesium chloride is decomposed, and magnesium hydroxide and hydrochloric acid are formed: $\text{MgCl}_2 + 2\text{H}_2\text{O} \rightleftharpoons \text{Mg}(\text{OH})_2 + 2\text{HCl}$; or possibly: $\text{MgCl}_2 + \text{H}_2\text{O} \rightleftharpoons \text{Mg}(\text{OH})\text{Cl} + \text{HCl}$, the free acid reddens blue litmus. The amount of hydrolysis in aqueous solutions of barium, strontium, and calcium chlorides is insignificant, although these salts, as well as the alkaline chlorides, are very slightly hydrolyzed. The hydrolysis of sodium chloride can be demonstrated by Emich's experiment:

Heat a little sodium chloride in a platinum crucible to bright redness, and add a couple of drops of water to the hot crucible so that the water assumes the spheroidal state. In a moment, transfer the water to a beaker containing a very faintly coloured solution of blue litmus—the litmus is reddened, showing the presence of an acid—hydrochloric acid. The salt remaining in the crucible is dissolved in water, and it turns red litmus blue, showing the presence of an alkali—sodium hydroxide.

Hydrolysis belongs to the type of “opposing reactions” previously discussed. Equilibrium occurs when the speed of hydrolysis is equal to the speed of re-combination of the products of the hydrolysis. If one or both the products of hydrolysis are removed from the sphere of the reaction, either by the formation of a precipitate, or by the liberation of a gas, the whole of the salt may be hydrolyzed. With bismuth chloride, BiCl_3 , for example, the basic salt $\text{BiCl}(\text{OH})_2$ is quantitatively precipitated: $\text{BiCl}_3 + 2\text{H}_2\text{O} \rightleftharpoons \text{BiCl}(\text{OH})_2 + 2\text{HCl}$. This reaction, in fact, is one of the standard methods employed in analysis for the precipitation of bismuth quantitatively from a solution.

In consequence of hydrolysis, when aqueous solutions of zinc or magnesium chlorides are concentrated by evaporation, some hydrogen chloride volatilizes, and a mixture of the oxide and chloride, or a basic compound— Mg_2OCl_2 , or $\text{MgO} \cdot \text{MgCl}_2$ —is formed. The anhydrous chloride cannot therefore be prepared by the evaporation of the aqueous solutions. If hydrogen chloride be prevented from leaving the solution by conducting the evaporation in a stream of hydrogen chloride, the hydrolysis of the magnesium chloride cannot proceed very far. Indeed, the excess of hydrogen chloride drives the hydrolysis backwards by increasing the concentration of hydrogen chloride in the system, so that it is possible to prepare anhydrous magnesium chloride by the evaporation of aqueous solutions in a stream of hydrogen chloride gas. See “zinc sulphate” for another view of hydrolysis.

Magnesium and zinc oxychlorides.—When magnesium chloride is heated in air, chlorine is evolved, and magnesium oxychloride, Mg_2OCl_2 ,

is formed; and at still higher temperatures, nearly all the chloride is decomposed and magnesium oxide, MgO , remains. See Weldon and Péchiney's process for chlorine, p. 277.

If calcined magnesia be made into a thick paste with a concentrated solution of magnesium chloride, or if zinc oxide and zinc chloride be treated in a similar manner, the mass hardens to a stone-like mass owing to the formation of an oxychloride. Hence the use of these mixtures as cements: *e.g.* zinc oxychloride is employed by dentists for stopping teeth; and magnesium oxychloride—*Sorel's cement*, or *magnesia cement*—is used in the manufacture of artificial stones, tiles, etc.

Solubility of the chlorides in water.—When some anhydrous chlorides are dissolved in water, heat is evolved, and the solution becomes warmer: *e.g.* $\text{CaCl}_2 + \text{Aq} = \text{CaCl}_{2\text{aq}} + 17.4 \text{ Cals.}$; with anhydrous barium chloride, 2.1 Cals. , and with anhydrous magnesium chloride, 436.0 Cals. On the contrary, when the crystalline hydrates are dissolved in water, heat is absorbed, and the solution becomes cooler: *e.g.* $\text{CaCl}_2 \cdot 6\text{H}_2\text{O} + \text{Aq} = \text{CaCl}_{2\text{aq}} - 4.3 \text{ Cals.}$; with crystalline barium chloride, $\text{BaCl}_2 \cdot 2\text{H}_2\text{O} - 4.9 \text{ Cals.}$, and with crystalline magnesium chloride, $\text{MgCl}_2 \cdot 6\text{H}_2\text{O} - 3.0 \text{ Cals.}$ Thus, the temperature can be reduced from 0° to -42° by a mixture of 2 parts of crystallized calcium chloride, with half its weight of snow or powdered ice; 3 parts of potassium chloride with 2 parts of snow will lower the temperature from 0° to -33° ; and with 10 parts of snow, from 0° to -11° ; 1 part of sodium chloride with 3 parts of snow will lower the temperature from 0° to -17.7° . Hence the use of such salts as **freezing mixtures**. Note the four methods of producing cold so far described: (1) rapid solution of some solids; (2) rapid evaporation of a volatile liquid; (3) rapid expansion of a compressed gas; and (4) cooling and compression of gases.

It is well to remember that the dissolution of a solid is attended by (1) the separation of the molecules against molecular attraction; (2) the liquefaction of the solid; and (3) combination of the solute with the solvent. The two former involve an expenditure of energy and exercise a cooling effect, the last alone usually exercises a heating effect; whether the dissolution of a solid will be an endo- or an exo-thermal process is thus determined by the relative magnitudes of these three effects. When liquids are dissolved, (2) will be zero; and with gases liquefaction will be attended by an evolution of heat.

An idea of the solubility of the chlorides of barium (BaCl_2), strontium (SrCl_2), calcium (CaCl_2), cadmium (CdCl_2), magnesium (MgCl_2), and zinc (ZnCl_2), will be obtained from the following table which represents the amount of salt, in grams, dissolved by 100 grams of water at the temperatures named:

TABLE XI.—SOLUBILITIES OF CHLORIDES.

Temperature.	BaCl_2	SrCl_2	CaCl_2	HgCl_2	CdCl_2	MgCl_2	ZnCl_2
0° . . .	32	43	59	3.5	9	53	208
50° . . .	44	72	126	11.6	10	60	470
100° . . .	59	102	155	38.0	14	73	615

The remarkable solubility of zinc chloride is worthy of special note. At 50° , for instance, zinc chloride forms a thick syrupy liquid with one-fourth of its weight of water. The great affinity of zinc and calcium chlorides (anhydrous) for water is utilized in chemical work. Calcium chloride is extensively employed in the laboratory for drying gases; and zinc chloride is a powerful dehydrating agent. It decomposes many organic compounds, taking from them the elements of water. Zinc chloride is also used as an antiseptic; in the preservation of wood from decay; and, owing to its power of dissolving metallic oxides, as a flux in soldering.

The smoothing of solubility curves.—The solubility curves of these salts would be misleading if plotted from the numbers indicated above, for many of these chlorides give solubility curves with a number of breaks which would thus be smoothed out and accordingly obscured. Care must always be exercised in smoothing data obtained at widely differing temperatures, because some important phenomena may thus be masked. The breaks correspond with the change in the solubility which attends the transformation of one hydrate into another as the temperature rises.

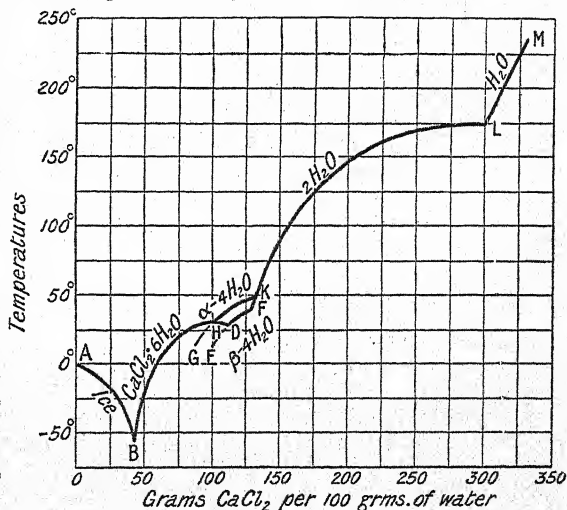


FIG. 115.—Solubility Curve of Calcium Chloride.

Calcium chloride may be taken to illustrate the principle, Fig. 115. At 29.8° the hexahydrate $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ changes into the α -tetrahydrate $\alpha\text{-CaCl}_2 \cdot 4\text{H}_2\text{O}$; at 45.3° the α -tetrahydrate passes into the dihydrate $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$; at 175.5° (not shown in the diagram) the dihydrate passes into the monohydrate $\text{CaCl}_2 \cdot \text{H}_2\text{O}$; and at 200° the monohydrate passes into the anhydrous salt CaCl_2 . The changes from one hydrate to the other produce the breaks indicated in Fig. 115. Each hydrate has its own particular solubility curve. There are two different tetrahydrates, each having its own specific solubility curve. They are distinguished as the α - and the β -tetrahydrates, the former being the stable hydrate whose solubility curve is shown in the diagram.

Deliquescence and efflorescence.—Each hydrate has its own specific vapour pressure at a particular temperature. The average vapour pressure of the water vapour in atmospheric air is equivalent to 8 or 9 mm. of mercury. If the vapour pressure of the hydrate be greater than the vapour

pressure of atmospheric moisture, the hydrate will *lose* water on exposure to the air—in other words, the salt will effloresce; on the contrary, if the vapour pressure of the hydrate be less than that of the atmospheric moisture, the salt will *absorb* moisture from the atmosphere, and deliquesce. Here are a few illustrations, taken at about 16° :

TABLE XII.—VAPOUR PRESSURES OF HYDRATES.

Salt.	Vapour pressure. mm.	Property.
$\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$	3.2	Deliquescent
$\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$	6.0	Deliquescent
$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$	27.8	Efflorescent
$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	24.2	Efflorescent

It is therefore possible to predict whether or not a salt will have a tendency to efflorescence or deliquescence when the vapour pressure of the hydrate is known. But a perfectly sound crystal of, say, sodium sulphate does not effloresce on exposure to the atmosphere. If, however, the change has commenced at any point, it will spread throughout the whole mass. This is in accord with the phase rule $F = C - P + 2$. In the perfect crystal there are two phases P , namely sodium sulphate, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$, and water vapour; and two components C , namely Na_2SO_4 and H_2O . Hence the system is bivariant ($F = 2$), so that the pressure of the water vapour and the temperature can vary arbitrarily within certain limits without altering the state of the system. If, however, some efflorescent salt be present, there will be three phases, and the system will be univariant ($F = 1$), so that for every temperature there is one and only one vapour pressure for equilibrium.

Manufacture of calcium and barium chlorides.—Calcium chloride is obtained as a by-product in many manufacturing operations—manufacture of potassium chlorate; of ammonia from ammonium chloride, etc. The salt crystallizes from its aqueous solutions in hexagonal crystalline prisms. Barium chloride is made by dissolving witherite, native barium carbonate, in hydrochloric acid. It is also manufactured from the mineral barytes, barium sulphate, BaSO_4 , by roasting the mineral with powdered coke, limestone, and calcium chloride. The barium sulphate is reduced to the sulphide, BaS , thus $\text{BaSO}_4 + 4\text{C} = \text{BaS} + 4\text{CO}$. The barium sulphide then reacts with the calcium chloride, forming calcium sulphide and barium chloride: $\text{BaS} + \text{CaCl}_2 = \text{CaS} + \text{BaCl}_2$. The mass is lixiviated with water; the calcium sulphide remains behind, and barium chloride passes into solution. When the solution is concentrated and cooled, barium chloride, $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$, separates in colourless rhombic plates.

Crystals of calcium oxychloride, $\text{Cl.Ca.O.Ca.OH} (+7\text{H}_2\text{O})$ are formed when lime is boiled with a concentrated solution of calcium chloride; and another basic chloride is formed when ammonia is made by heating ammonium chloride with lime. The oxychlorides do not all have the same composition.

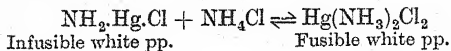
§ 9. Mercury, Silver, Copper, and Gold Chlorides.

Mercury, silver, copper, and gold form a series of chlorides— HgCl , CuCl , AgCl , AuCl —similar in type to the alkaline chlorides— KCl , NaCl —but they are all very sparingly soluble in water. In these chlorides the metal appears to behave as a monad; but, in addition, mercury and copper respectively form HgCl_2 and CuCl_2 , in which the metals behave as dyads; gold also forms a second chloride— AuCl_3 —in which the metal behaves as a triad. The lower chlorides are named: mercurous, cuprous, and aurous chlorides; and the higher chlorides are called: mercuric, cupric, and auric chlorides.

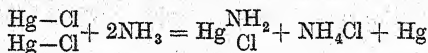
Mercuric chloride, HgCl_2 —also called *corrosive sublimate*—is made on a commercial scale by heating a mixture of mercuric sulphate with sodium chloride: $2\text{NaCl} + \text{HgSO}_4 = \text{Na}_2\text{SO}_4 + \text{HgCl}_2$; a little manganese dioxide, MnO_2 , is usually added with the object of preventing the formation of the lower chloride. Mercuric chloride sublimes as a white translucent mass. The salt can be obtained in needle-like rhombic prisms by cooling a solution saturated with the salt at 100° . There is a great difference in the solubility of the salt at 100° and at lower temperatures; thus 100 grams of solution at 0° contains 3.5 grams of salt; and at 100° , 38 grams. It is more soluble in mercury. The salt melts at 277° and vaporizes about 301° . Like magnesium and zinc chlorides, mercuric chloride readily forms oxychlorides and also double salts, e.g. $\text{HgCl}_2 \cdot \text{HCl}$; $\text{HgCl}_2 \cdot \text{HgO}$; and the *sal alembroth*, or salt of wisdom, obtained by the early chemists by crystallization from a solution of equal parts of ammonium and mercuric chlorides. The double salts are more soluble than mercuric chloride, and they are much used in making antiseptic solutions for taxidermists, etc. Mercuric chloride is a valuable antiseptic, and is used for washing wounds, etc., in surgery. Dilute solutions, 1 part of salt per 1000 parts of water, or per 2000 parts of water, are usually employed. Mercuric chloride is a virulent poison. The antidote is albumen (the white of a raw egg). This forms with the salt an insoluble mass which can be removed from the stomach.

When a solution of mercuric chloride is treated with an excess of aq. ammonia, a white bulky precipitate of mercuric amido chloride, $\text{NH}_2 \cdot \text{HgCl}$, or *infusible white precipitate* or simply *white precipitate* is formed; if the order of mixing be reversed, and particularly if mercuric chloride be added to a boiling solution of ammonium chloride, mercuric diammino chloride, $\text{Hg}(\text{NH}_2)_2\text{Cl}_2$, is formed. It is called *fusible white precipitate* because, when heated, it fuses and then volatilizes, whereas infusible white precipitate volatilizes without fusing. Infusible white precipitate is used in pharmacy for making ointments. When infusible white precipitate is treated with hydrochloric acid, the reaction is symbolized: $\text{NH}_2 \cdot \text{HgCl} + 2\text{HCl} = \text{NH}_4\text{Cl} + \text{HgCl}_2$; and with hydrofluoric acid, mercuric chloride and fluoride and ammonium chloride are formed. This has been taken to indicate that the chlorine atom in the original compound is attached to the mercury atom: $\text{NH}_2 \cdot \text{HgCl}$, and not to the nitrogen atom: $\text{Cl} - \text{N} \leq \overset{\text{Hg}}{\text{H}_2}$. This is confirmed by the production of the amido chloride by the action of sodamide on mercuric chloride: $\text{NaNH}_2 + \text{HgCl}_2$

= NaCl + NH₂.Hg.Cl; by the formation of a yellow crystalline precipitate, HgPtCl₆(NH₂), when it is treated with chloroplatinic acid, H₂PtCl₆; and by the formation of an ethyl derivative NH(C₂H₅).Hg.Cl. The relation between fusible and infusible white precipitates as shown by the reversible reaction:



Mercurous chloride, HgCl or Hg₂Cl₂—also called *calomel*. If mercuric chloride be heated along with metallic mercury, a sublimate of mercurous chloride is obtained as a white powder. The mixture used for the preparation of mercuric chloride can also be employed if the manganese oxide be omitted, and metallic mercury introduced: HgCl₂ + Hg = 2HgCl. Mercurous chloride can also be made by the direct union of mercury and chlorine; and by the addition of a soluble chloride, or, better, hydrochloric acid, to a soluble mercurous salt. Mercurous chloride is precipitated. One litre of water dissolves but 0.002 gram of mercurous chloride at 18°. Mercurous chloride sublimes at 373°, and it slowly blackens when exposed to light. This is said to be due to the reversal of the reaction just indicated. The blackening is thus due to the separation of mercury, 2HgCl = HgCl₂ + Hg. The salt is also blackened by contact with alkalis and aqueous ammonia—hence the term “calomel” from the Greek *καλομέλας* (kalomelas), black. This black powder appears to be mercuric amido-chloride, Hg(NH₂)Cl, mixed with metallic mercury, and the reaction is represented:



Mercurous chloride absorbs dry ammonia gas forming Hg₂Cl₂.2NH₃. Mercurous chloride reacts with alkali hydroxides as symbolized: 2HgCl + 2KOH = Hg₂O + H₂O + 2KCl; and with alkali carbonates: 2HgCl + Na₂CO₃ = Hg₂CO₃ + 2NaCl. Mercurous chloride occurs native in Spain and some other countries. The mineral is called *horn quicksilver*. Both mercurous and mercuric chlorides are employed in medicine.

The vapour density of mercurous chloride.—The vapour density of mercurous chloride at 518° is 235.5 (O₂ = 32). This number agrees with 235.5 required for the simple molecule HgCl. If mercury be a dyad, and if it has a constant unchangeable valency, the formula should be Hg₂Cl₂, that is, Cl—Hg—Hg—Cl. In that case, the observed vapour density means that the mercurous chloride dissociates into mercuric chloride and mercury: Hg₂Cl₂ = Hg + HgCl₂; the theoretical vapour density of which is $\frac{1}{2}(200 + 270.9) = 235.5$, the same as for HgCl. The vapour density determination does not therefore furnish a definite answer to the obvious question. The vapour of calomel forms an amalgam with gold when a piece of gold leaf is suspended in the vapour, owing to the presence of mercury vapour. It is possible that gold leaf decomposes calomel vapour, and, at best, the experiment only shows that *some* dissociation occurs, whereas the observed density requires *complete* dissociation. The argument based on the constant valency of mercury is of little value. Hence an *experimentum crucis*—that is, an experiment which will decide the question without ambiguity—is still wanting. The best arguments in

favour of the dissociation hypothesis ($\text{Hg}_2\text{Cl}_2 \rightleftharpoons \text{HgCl}_2 + \text{Hg}$) are due (1) to H. B. Baker. He claims that the vapour density of calomel, perfectly dry, corresponds with the formula Hg_2Cl_2 ; but if the calomel be not so completely dried, it is probable that Hg_2Cl_2 dissociates into Hg and HgCl_2 , analogous with the behaviour of perfectly dried ammonium chloride, which is known to behave in this manner. (2) A. Smith and A. W. C. Menzies (1910) showed that the vapour pressure of a mixture of calomel and mercury will be equal to the sum of the vapour pressure of mercury and calomel if there be no dissociation, and less than this sum if dissociation of the calomel vapour occurs. Experiment showed that the observed vapour pressure of the mixture is less than that calculated for complete dissociation. This was traced to a slight lowering of the vapour pressure of mercury owing to its dissolving a little calomel. Allowing for this, Smith and Menzies' measurements give numbers in close agreement with the view that the vapour of calomel is completely dissociated.

Cupric chloride, CuCl_2 .—This salt can be made by dissolving the metal in a mixture of nitric and hydrochloric acids, as well as by the usual methods: burning the metal in chlorine, etc. Cupric chloride, when anhydrous, is a dark brown solid. Free chlorine can be detected in the vapour of cupric chloride at 344° ; and at about 400° it furnishes cuprous chloride and chlorine gas. Cupric chloride dissolves in a small quantity of water, forming a deep green solution which deposits rhombic crystals of $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$ on evaporation. If the aqueous solution of cupric chloride be diluted with a larger volume of water, the solution becomes blue and loses its green tint; the green tint is restored if the blue solution be concentrated by evaporation. When ammonia gas is passed into an aqueous solution of cupric chloride, blue crystals of $\text{CuCl}_2 \cdot 4\text{NH}_3 \cdot \text{H}_2\text{O}$ are deposited. This compound loses all its water and half its ammonia when heated between 120° – 125° . When the anhydrous chloride is exposed to ammonia gas, a blue solid, $\text{CuCl}_2 \cdot 6\text{NH}_3$, is formed; it is insoluble in liquid ammonia, but soluble in water, and gradually dissociates into the tetrammine, $\text{CuCl}_2 \cdot 4\text{NH}_3$, the diammino, $\text{CuCl}_2 \cdot 2\text{NH}_3$; and at higher temperatures this compound decomposes into nitrogen, cuprous chloride, ammonia, and ammonium chloride: $6(\text{CuCl}_2 \cdot 2\text{NH}_3) = 6\text{CuCl} + 6\text{NH}_4\text{Cl} + 4\text{NH}_3 + \text{N}_2$. The constitution of the "metal-ammonia" compounds, called the "ammines," is discussed later.

Cuprous chloride, CuCl .—When a solution of cupric chloride in concentrated hydrochloric acid is digested with metallic copper, and the acid solution is poured into water, a white precipitate of cuprous chloride separates. When either cuprous or mercurous chloride is heated in a current of chlorine, the "-ic" salt is produced, e.g. $2\text{CuCl} + \text{Cl}_2 = 2\text{CuCl}_2$; and conversely, when either cupric or mercuric chloride is heated in a current of hydrogen, or brought in contact with nascent hydrogen, the "-ous" salt is produced, followed *immediately* by the reduction of the chloride to the metal. Cuprous chloride melts between 415° and 422° . Stannous chloride, SnCl_2 , reduces mercuric to mercurous chloride, and the latter, in turn, to metallic mercury; sulphurous acid, H_2SO_3 , reduces cupric to cuprous chloride. Conversely, cuprous chloride can be oxidized to cupric chloride by exposing the hydrochloric acid solution of cuprous chloride to the atmosphere. The solution becomes brown and then deposits a greenish-blue solid, $\text{CuCl}_2 \cdot 3\text{CuO} \cdot 3\text{H}_2\text{O}$, which appears to resemble

the mineral *atacamite*. A bromide analogue is also known. Cuprous chloride, dissolved in hydrochloric acid or in ammonia, readily absorbs carbon monoxide. Both solutions are employed in gas analysis for the removal of carbon monoxide from gaseous mixtures. The amount of gas absorbed never exceeds the ratio $\text{Cu} : \text{CO}$, and it is probable that a compound $\text{CuCl} \cdot \text{CO} \cdot 2\text{H}_2\text{O}$ is formed. When a solution approaching saturation is brought in contact with a gas free from carbon monoxide, some carbon monoxide may leave the solution and contaminate the gas.

Silver chloride, AgCl .—Silver chloride occurs in nature as the mineral *horn silver*. Like mercurous chloride, silver chloride is formed when a soluble chloride is added to a solution of a silver salt. The white precipitate melts between 451° and 455° , forming a yellow liquid which becomes darker the higher the temperature. The liquid solidifies on cooling to a tough horny mass. Silver chloride absorbs ammonia gas, forming $\text{AgCl} \cdot 2\text{NH}_3$; and $\text{AgCl} \cdot 3\text{NH}_3$, some properties of which will be discussed later. Silver chloride is soluble in concentrated hydrochloric acid, in alkaline chlorides, aqueous ammonia, potassium cyanide, and in sodium thiosulphate. Its solubility in water was indicated on p. 190. See also p. 323.

Action of light on silver chloride.—Thoroughly dried and purified silver chloride can be exposed to sunlight for several hours without darkening; but if a trace of moisture be present, the chloride darkens, assumes a violet tint, which passes into brown, and finally turns black. It is probable that the chloride decomposes into free chlorine and silver subchloride, Ag_2Cl ; thus: $4\text{AgCl} \rightleftharpoons 2\text{Ag}_2\text{Cl} + \text{Cl}_2$. The system is in equilibrium when the vapour pressure of the chlorine has reached a certain value. Some hold that the molecule of the subchloride should be represented Ag_4Cl_2 , or $\text{Ag}_2 = \text{Cl} - \text{Cl} = \text{Ag}_2$. When the silver subchloride is exposed to the action of chlorine gas in darkness, silver chloride, AgCl , is formed. The amount of silver chloride decomposed, for equilibrium, depends upon the intensity of the light. If chlorine be removed from the sphere of the reaction, by, say, mixing the silver salt with a suitable organic compound which "binds" the chlorine, the reaction proceeds to an end. These properties of silver chloride are employed in photography. Silver bromide and silver iodide appear to behave in an analogous manner, but the bromide is generally considered most suitable for photographic purposes.

Photography.—A celluloid film or glass plate is coated on one side with a film of gelatine containing, say, silver bromide in suspension, and dried. The plate is placed in the camera, and *exposed* by focusing the image of the object to be photographed on to the plate for a moment. The silver bromide is affected in some unknown way so that the most intense change occurs where the light is brightest, while the change is less intense in the shadows. No visible change is apparent until the plate is *developed*. Chemists who have specialized in the study of the so-called *latent image* are not agreed as to what the exposure does to the plate so that after development an image of the object appears. The rivals are the sub-halide hypothesis, the reduced silver hypothesis, the solid solution hypothesis, etc. The plate is developed by treating it with a reducing agent—ferrous oxalate, pyrogallol, or some special developer. The developer continues the change started by the light, but it is without action on the unexposed parts of the plate. As a result, finely divided silver is deposited on the

parts of the plate illuminated by the light reflected from the object. The deposit is thickest where the light was most intense. Hence, the dark parts of the object appear lightest on the plate, and the light parts dark. The image is thus the reverse of the object, and the plate is accordingly called a *negative*. The silver salt which has not been affected by the light nor by the developer is now removed, and the image thus *fixed* on the plate by immersing the plate in a solution of sodium thiosulphate. The plate is then washed and dried. A *print* is made by laying the negative upon sensitized printing paper—that is, paper prepared in a similar way to the negative—so that the light must pass through the negative before striking the paper. The negative absorbs the light in proportion to the thickness of the deposit of silver, so that the print has the same shading as the object. The paper is then treated with a solution of sodium thiosulphate to fix the image. The print may be *toned* by immersing it in a solution of gold chloride so that some of the silver is replaced by gold; this gives the print a warm reddish tone; if a platinum salt be used instead of gold, a steel-grey tone is produced. The image on the print will be the reverse of that on the negative, and will therefore correspond with the object. Hence the print may be called the *positive*.

Auric chloride, AuCl_3 .—If gold be dissolved in a mixture of hydrochloric and nitric acids (*aqua regia*), yellow needle-like crystals of a complex acid—hydrochloroauric acid, HAuCl_4 —separate from moderately concentrated solutions. If the solution be evaporated to dryness, and the residue dissolved in a little water, reddish crystals of $\text{AuCl}_3 \cdot 2\text{H}_2\text{O}$ can be obtained by concentrating the solution. These crystals lose water and give brown deliquescent crystals of the anhydrous chloride when heated to a low temperature. The crystals contain traces of hydrochloric acid, and slowly decompose at 100° ; $2\text{AuCl}_3 = 2\text{Au} + 3\text{Cl}_2$. Auric chloride is formed as a dark red crystalline mass—mixed with a little metallic gold and aurous chloride—when a stream of dry chlorine is passed over metallic gold at 200° – 220° . Hydrochloroauric acid, HAuCl_4 , forms a series of complex salts with the alkaline chlorides: $\text{NaAuCl}_4 \cdot 2\text{H}_2\text{O}$, etc. These salts are called *chloroaurates*. The univalent radicle is " AuCl_4 ." The constitution of complex salts of this kind is discussed later.

Aurous chloride, AuCl .—When auric chloride, free from traces of hydrogen chloride, is heated between 170° and 180° , it passes into a yellowish-white powder which is aurous chloride; below this range of temperature the action is very slow, and above, the yield is small. The admixed auric chloride can be removed by washing the product with thoroughly dried ether. Aurous chloride in contact with water decomposes into auric chloride and gold, even at ordinary temperatures: $3\text{AuCl} = \text{AuCl}_3 + 2\text{Au}$. If aurous oxide, Au_2O , be dissolved in hydrochloric acid, it forms auric chloride and metallic gold. Both the gold chlorides are unstable, and decompose when heated above 180° into chlorine and the metal. Dry chlorine unites readily with aurous chloride at ordinary temperatures, forming auric chloride. Gold is precipitated from its solutions of the chloride by reducing agents of all kinds—stannous chloride, ferrous sulphate, formic, sulphurous, and oxalic acid, etc.

§ 10. Tin and Lead Chlorides.

Stannic chloride, SnCl_4 .—This salt can be made by heating powdered tin with an excess of mercuric chloride, $2\text{HgCl}_2 + \text{Sn} = 2\text{Hg} + \text{SnCl}_4$; or by the following process:

Place 60 grams of metallic tin in a flask, and heat it by placing a Bunsen's burner underneath. When the tin is melted, pass a stream of dry chlorine, prepared as indicated above, so that the gas plays directly on the metal, Fig. 116. When the metal in the flask has all disappeared, add some tinfoil to the liquid with the receiver, and close the receiver with a stopper. After standing 12 hours so as to allow the dissolved chlorine to react with the tin, distil the product from a retort, and reject the first few drops of the distillate. If the later distillate be coloured, repeat the treatment with tinfoil, etc., and collect the fraction which distils at 113.9° .

Stannic chloride is a fuming colourless liquid which boils at 113.9° . The liquid was known to A. Libavius in 1605, and was called *spiritus fumans Libavii*. It unites with water, forming a series of crystalline hydrates: $3\text{H}_2\text{O}$; $5\text{H}_2\text{O}$; and $8\text{H}_2\text{O}$. The pentahydrate $\text{SnCl}_4 \cdot 5\text{H}_2\text{O}$ is used as a mordant¹ in dyeing under the name "oxymuriate of tin," and

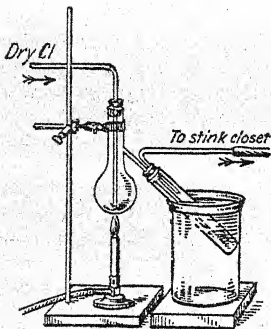


FIG. 116.—Preparation of Stannic Chloride.

also incorrectly as *pink salt*. Aqueous solutions of stannic chloride are not stable owing to hydrolysis: $\text{SnCl}_4 + 4\text{H}_2\text{O} \rightleftharpoons 4\text{HCl} + \text{Sn}(\text{OH})_4$. Stannic chloride forms a series of complex salts: potassium chlorostannate, $\text{SnCl}_4 \cdot 2\text{KCl}$, or K_2SnCl_6 , in which the radicle " SnCl_6 " is bivalent. The corresponding ammonium chlorostannate, $(\text{NH}_4)_2\text{SnCl}_6$, has been much used in dyeing under the name *pink salt*. The concentrated aqueous solution does not decompose when heated, but the dilute solution furnishes metastannic acid. Potassium fluostannate, K_2SnF_6 , analogous to the chlorostannate, is obtained by the action of hydrofluoric acid on potassium stannate. The aqueous solution is acid, and if an equivalent amount of potassium

hydroxide be added to the solution of the fluostannate, the liquid becomes neutral, and lamellar crystals of potassium hydroxy fluostannate, $\text{K}_2\text{SnF}_5(\text{OH}) \cdot \text{H}_2\text{O}$, are deposited.

Stannous chloride, SnCl_2 .—Dissolve metallic tin in hydrochloric acid, and evaporate the solution until monoclinic crystals of the so-called *tin salt*, $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$, separate. This hydrate loses its water and forms anhydrous stannous chloride when dried *in vacuo*. The anhydrous salt is also made by heating metallic tin in a stream of hydrogen chloride; and by heating a mixture of metallic tin with mercuric chloride: $\text{HgCl}_2 + \text{Sn} = \text{SnCl}_2 + \text{Hg}$. The mercury volatilizes and leaves a residue of stannous chloride. Stannous chloride is soluble in a small amount of water; the addition of an excess of water or exposure to the air leads to the precipitation of a basic chloride: $2\text{SnCl}_2 + 2\text{H}_2\text{O} = \text{SnCl}_2 \cdot \text{SnO} \cdot \text{H}_2\text{O}$

¹ A mordant is a substance which can be precipitated in the fibre of a fabric to be dyed, and which then combines with the dye to form an insoluble compound.

+ 2HCl. The formula of stannous oxychloride is variously represented: $\text{SnCl}_2\text{SnO} \cdot \text{H}_2\text{O}$; $\text{Sn}_2\text{OCl}_2 \cdot \text{H}_2\text{O}$; $\text{SnO} \cdot \text{HCl}$; $\text{Sn}(\text{OH})\text{Cl}$. Stannous chloride is a powerful reducing agent, and readily combines with oxygen or with chlorine. Thus, ferric salts are reduced to the ferrous state: $2\text{FeCl}_3 + \text{SnCl}_2 = \text{SnCl}_4 + 2\text{FeCl}_2$; mercuric salts are reduced to mercurous salts, and finally to mercury. When solutions of stannous chloride are exposed to the air, oxygen is absorbed, and stannous oxychloride and stannic chloride are formed: $6\text{SnCl}_2 + \text{O}_2 + 2\text{H}_2\text{O} = 4\text{Sn}(\text{OH})\text{Cl} + 2\text{SnCl}_4$. Stannous chloride boils at 606° . The vapour density of stannic chloride, above 900° , corresponds with $\text{SnCl}_2 + \text{Cl}_2$, and at lower temperatures with the mixture: $\text{SnCl}_4 \rightleftharpoons \text{SnCl}_2 + \text{Cl}_2$. Stannous chloride also combines with iodine, forming SnCl_2I_2 , and it is used for titrating iodine in volumetric analysis.

Lead chloride, PbCl_2 .—This salt separates as a white curdy precipitate when hydrochloric acid or a soluble chloride is added to a solution of a lead salt. It is also made by dissolving lead oxide or carbonate in hot hydrochloric acid. On cooling, the solution deposits crystals of lead chloride, PbCl_2 , which melt at 495° . The salt is generally said to be fairly soluble in hot water, and sparingly soluble in cold water: 100 grams of water at 0° dissolve 0.67 gram of PbCl_2 ; at 50° , 1.70 grams; and at 100° , 3.34 grams. When lead chloride is heated in air, lead oxychloride, Pb_2OCl_2 , or $\text{PbCl}_2 \cdot \text{PbO}$, is formed. By adding hot limewater to a boiling solution of lead chloride, $\text{PbCl}_2 \cdot \text{PbO} \cdot \text{H}_2\text{O}$, or $\text{Pb}(\text{OH})\text{Cl}$, separates. This compound is used as a white pigment under the commercial name "Pattinson's white lead," and the pigment "Cassel's yellow" is a mixture containing one or more oxychlorides of lead, approximately $7\text{PbO} + \text{PbCl}_2$, made by heating lead oxide with ammonium chloride.

Lead tetrachloride, PbCl_4 .—This is an unstable chloride formed when chlorine is passed through hydrochloric acid in which lead chloride is suspended; or when lead dioxide is dissolved in concentrated hydrochloric acid. The salt, $\text{PbCl}_4 \cdot 2\text{NH}_4\text{Cl}$, or $(\text{NH}_4)_2\text{PbCl}_6$, is precipitated when ammonium chloride is added to the solution. This salt ammonium chloroplumbate is said to be isomorphous with ammonium chlorostannate. If ammonium chloroplumbate be treated with concentrated sulphuric acid in the cold, a yellow oil, PbCl_4 , is obtained of sp. gr. 3.18 at 0° . The tetrachloride is unstable, decomposing: $\text{PbCl}_4 = \text{PbCl}_2 + \text{Cl}_2$. Lead tetrachloride freezes to a yellow crystalline solid at about -15° , and when treated with hydrochloric acid forms yellow crystals of presumably hydrochloroplumbic acid, H_2PbCl_4 ; and when treated with a little water furnishes what is thought to be lead tetrahydroxide, or plumbic acid, H_4PbO_4 . This compound is interesting because it shows that very probably lead can be quadri- as well as bi-valent.

§ 11. Iron, Aluminium, Chromium, Manganese, Cobalt, and Nickel Chlorides.

Ferric chloride, FeCl_3 .—The anhydrous chlorides of the metals can generally be made by the action of chlorine upon the metals, or by the action of chlorine, or the vapours of carbonyl chloride, or carbon tetrachloride upon the metallic oxides at a low red heat. The first-named operation can be conducted in the following manner:

A bundle of iron wire is placed in the middle of a hard glass tube *A*, Fig. 117. Connect one end of this tube with an apparatus for generating chlorine and two sulphuric acid wash-bottles for drying the gas; connect the other end of the tube with a dry receiver, made from a wide-necked bottle, with a two-holed stopper, and one tube leading to the sink closet as indicated in the diagram. Conduct a fairly rapid stream of chlorine through the tube, and when all the air has been expelled, gently warm the tube with a flame which does not touch the glass. In a short time, the iron and chlorine will react with brilliant sparks. The ferric chloride which is formed can be sublimed into the receiver by warming the neck, etc., with a second large flame moving to and fro along the tube. When the iron is all converted into ferric chloride, shake the salt into a dry wide-mouthed bottle. The bottle must be well-stoppered—stopper *S*, Fig. 117—because the salt is very hygroscopic or deliquescent.

Ferric chloride forms hexagonal crystals which appear greenish by reflected light, and dark red by transmitted light. By heating the anhydrous chloride and bromine in a sealed tube at 100°, dark red crystals

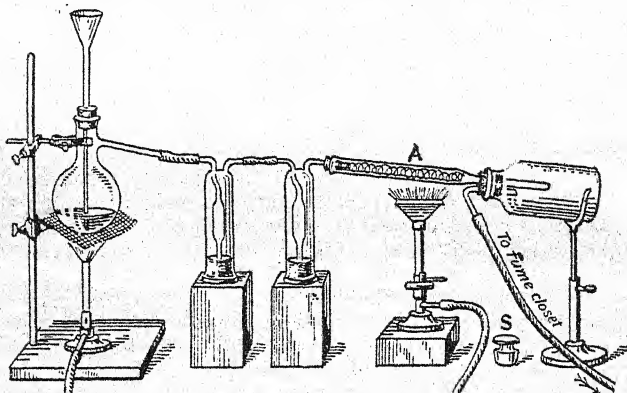


FIG. 117.—Preparation of Anhydrous Ferric Chloride.

of hygroscopic *ferric bromochloride*, FeCl_2Br , are formed. Anhydrous ferric chloride dissolves in water with the evolution of much heat. When aqueous solutions are evaporated, crystals of $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$ are formed as the solution cools, but the anhydrous salt cannot be obtained by the evaporation of aqueous solutions on account of hydrolysis. The hydrated salt can be obtained by dissolving iron, iron carbonate, or iron oxide in hydrochloric acid, and oxidizing the solution with a little nitric acid, or hydrogen peroxide. Hydrolysis occurs when aqueous solutions of ferric chloride are boiled or left standing some time. An insoluble oxychloride, or a soluble hydroxide and hydrochloric acid are formed. The two latter can be separated by dialysis.

Dialysis.—While studying the rate of diffusion of salts through membranes of parchment paper, Thomas Graham (1861) noticed that certain substances diffuse very slowly through the parchment. Gelatine and glue might be cited in illustration. The membrane is nearly impervious to these substances. Graham applied the term *colloids*—from the Greek *κόλλα* (*kolla*), glue—to those substances which diffuse but slowly through the parchment. On the other hand, substances like sodium chloride,

magnesium sulphate pass through the membrane much more quickly. Crystalline salts are typical of those substances which diffuse rapidly, and hence Graham called them *crystalloids*. In illustration :

CRYSTALLOIDS.

Potassium chloride.
Cane sugar.
Magnesium sulphate.
Hydrochloric acid.
Sodium chloride.
Barium chloride.

COLLOIDS.

Albumen.
Gums.
Starch.
Gelatinous aluminium hydroxide.
Gelatinous ferric hydroxide.
Gelatinous silicic acids.

It must be added that a great many substances can exist in both the colloidal and crystalloid condition, so that it is better to distinguish between the colloidal and crystalline conditions of matter rather than between "crystalloids" and "colloids."¹ In modern chemistry the term "colloid" comprises much of what was formerly called "amorphous," although some so-called colloids may after all be crystalline. It must not be supposed that the colloids do not pass through the parchment at all. Graham found that when the time of diffusion of hydrochloric acid—HCl—was taken as unity, the rate of diffusion of an equal quantity of sodium chloride was 2.3, cane sugar 7, egg albumen 49, and caramel 98. On account of these great differences, Graham proposed a useful method of separating colloids from crystalloid substances in solution. The crystalloid is removed by diffusion through a membrane of parchment, bladder, or some similar substance.

The process is called *dialysis*—from the Greek *διά* (dia), through; *λύω* (lyo), I loosen. The operation will be understood from the following description :



FIG. 118.—Dialyzer.

A piece of parchment or bladder is bound across one end of a glass or gutta-percha hoop so as to form a kind of shallow dish, Fig. 118, narrower at the base than the open top. A mixed solution of albumen (the white of an egg) and potassium chloride in water is poured into the "dish." This vessel is placed in another dish *B*, containing distilled water. The water in the outer vessel is renewed every few hours. The dish containing the mixed solution is covered by a clock glass to protect it from dust. In about three days, practically all the potassium chloride will have passed through the membrane into the outer vessel, while the egg albumen will remain in the inner compartment. The whole apparatus is called a dialyzer. See also Fig. 329.

If a few drops of ammonia be added to a solution of ferric chloride, a reddish-brown precipitate of ferric hydroxide will be formed. This redissolves in the unchanged ferric chloride. When the solution of ferric chloride is saturated with the ferric hydroxide, any further addition of ammonia will give a permanent precipitate of ferric hydroxide. Add a few drops of hydrochloric acid to dissolve the precipitated ferric hydroxide, and dilute the solution with water so that it contains about 5 per cent. of solid in solution. Pour this solution into the dialyzer, the soluble ferric chloride and ammonium chloride pass into the outer vessel, and a dark red liquid, called "dialyzed iron," remains in the inner vessel. If dialyzed

¹ Wo. Ostwald (1911) advocates the term *dispersoid system* in place of "colloidal solution."

iron be allowed to stand for some weeks in a glass vessel, it gelatinizes. Here, then, we have two kinds of colloids: *fluid* colloids are said to be in the sol condition; and *gelatinous* or pectinous colloids are said to be in the gel condition. When the solvent is water, the colloids are either hydrogels (gelatinous) or hydrosols (fluid); if alcohol be the solvent, alcogels or alcosals, etc. The dialyzed iron is in the hydrosol condition.

Solutions of the salts of the alkalies and alkaline earths can be readily obtained optically empty, but solutions of salts of iron, copper, aluminium, chromium, etc., cannot be so prepared. The explanation turned on the hydrolysis of the last-named salts whereby colloidal hydroxides are formed, that this is responsible for the turbidity. The hydroxides of the alkalies and alkaline earths are not colloidal, and hence if hydrolysis of their salts did occur, the liquid would remain optically clear. The hydrolysis of these salts is neatly illustrated by B. L. Vanzetti's experiment:

A test-tube is filled, approximately three-fourths its height, with a 5–10 per cent. solution of gelatin containing faintly alkaline phenolphthalein. When the gelatin has solidified, a layer of a 10 per cent. aqueous solution of ferric chloride is poured over it. As the salt diffuses into the gelatin two strata appear; the lower stratum is colourless, the upper one opaque; the decolorized stratum is produced by the acid which diffuses into the gelatin faster than the ferric hydroxide which makes the opaque stratum. Other coloured salts—e.g. nickel chloride or copper sulphite—can be used.

The result of many investigations on the hydrolysis of ferric salts is to show that the products of hydrolysis are not constant in composition. The main reaction is probably $\text{FeCl}_3 + 3\text{H}_2\text{O} \rightleftharpoons 3\text{HCl} + \text{Fe}(\text{OH})_3$ where the amount of hydroxide formed depends on the concentration and temperature of the solution; a secondary reaction $\text{FeCl}_3 + \text{Fe}(\text{OH})_3 \rightleftharpoons \text{Fe}_2\text{O}_3 + 3\text{HCl}$ is also supposed to occur since the neutralization of the hydrochloric acid by ammonia furnishes some ferric oxide Fe_2O_3 . The colloidal hydroxide formed is supposed to be a solid solution of $\text{FeCl}_3 \cdot n\text{Fe}(\text{OH})_3$ where n is very large.

In an earlier chapter, solutions were defined to be "mixtures which appear clear and homogeneous in ordinary daylight, and which cannot be separated into their constituent parts by filtration through paper, and by decantation." It is now possible to apply Graham's dialysis test, and subdivide solutions¹ according as the substance in solution diffuses rapidly—crystalloid solutions; or slowly—colloidal solutions—through parchment paper. To summarize:

Solutions are clear and homogeneous in daylight.

1. *Crystalloid solutions* are optically inert, and the dissolved matter diffuses rapidly through parchment.
2. *Colloidal solutions* give an opalescence with Tyndall's test, they usually appear heterogeneous under the ultramicroscope, and the substance in solution diffuses very slowly through parchment.
 - (a) *Sol*. The fluid colloidal condition.
 - (b) *Gel*. The solid or gelatinous colloidal condition.

It must be clearly understood that no hard-and-fast line of demarcation can be made between mere suspensions, colloidal and crystalloid solutions. Perfect definitions are not yet possible.

¹ The criteria "homogeneous in daylight" and "speed of diffusion" are quite arbitrary. The same remark applies to "filtration through paper," because colloidal solutions could be separated into their constituents if a compact enough filtering medium—say, more or less vitreous unglazed earthenware—were employed.

The vapour density of ferric chloride.—Anhydrous ferric chloride boils about 280° . At temperatures below 400° , the vapour density corresponds with Fe_2Cl_6 ; above that temperature the vapour density diminishes, until, at about 750° , the vapour density corresponds with the molecule FeCl_3 ; possibly also some dissociation,

$\text{Fe}_2\text{Cl}_6 \rightleftharpoons 2\text{FeCl}_3 + \text{Cl}_2$, occurs; at any rate, free chlorine can be detected in the vapour at 122° – 123° . The vapour density curve at different temperatures is indicated in Fig. 119. The effect of ferric chloride on the

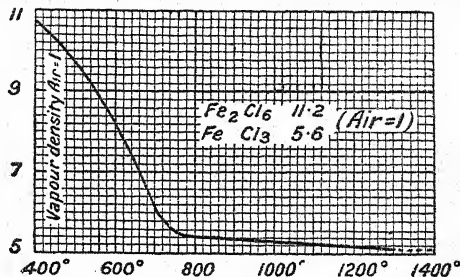


FIG. 119.—Vapour Density of Ferric Chloride.

boiling point of ether or alcohol corresponds with the formula FeCl_3 .

The hydrates of ferric chloride.—The curve *AB*, Fig. 120, represents

the effect of additions of anhydrous ferric chloride on the freezing point of solutions of water. The freezing point falls rapidly with the separation of ice from the solution until the eutectic point *B* at -55° is attained, and the solution contains 2.75 molecules of Fe_2Cl_6 per 100 molecules of the solvent. At the eutectic point, the whole solution freezes *en masse*. The solid is a mixture of ice and

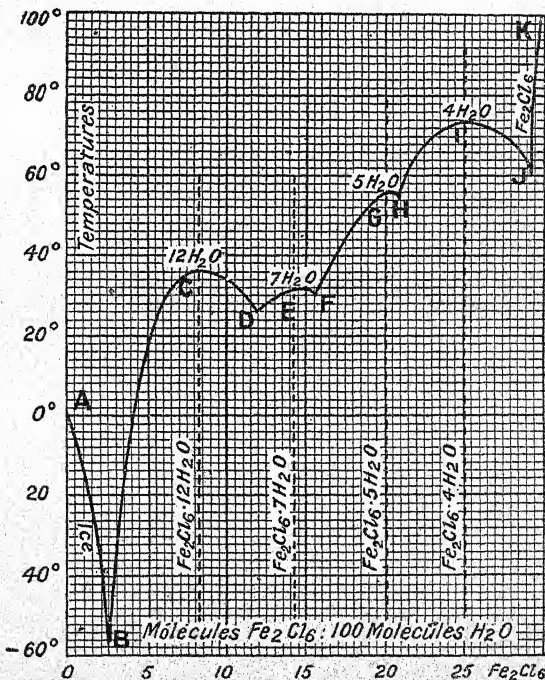


FIG. 120.—Freezing Points of the Hydrates of Ferric Chloride.

$\text{Fe}_2\text{Cl}_6 \cdot 12\text{H}_2\text{O}$

Any further addition of ferric chloride raises the temperature at which solid separates from the solution, and the solid which separates is the dodecahydrate, $\text{Fe}_2\text{Cl}_6 \cdot 12\text{H}_2\text{O}$, alone. The curve *BC* thus represents the solubility of the dodecahydrate

in water. This salt continues to separate until a temperature of 37° is attained. The solution is then nothing but fused $\text{Fe}_2\text{Cl}_6 \cdot 12\text{H}_2\text{O}$. This hydrate melts at 37° . Any further addition of ferric chloride to the solution depresses the temperature at which solid dodecahydrate separates until a second eutectic D is reached at 27.4° . The solution then solidifies *en masse*. The eutectic solid is a mixture of the dodecahydrate and the heptahydrate: $\text{Fe}_2\text{Cl}_6 \cdot 7\text{H}_2\text{O}$. By adding more salt, the temperature of solidification is raised; and the heptahydrate separates from the solution. This continues until a temperature of 32.5° is attained, when the solution is virtually fused heptahydrate: $\text{Fe}_2\text{Cl}_6 \cdot 7\text{H}_2\text{O}$. Any further addition of ferric chloride depresses the temperature at which the heptahydrate separates until a third eutectic is attained. The eutectic mixture of heptahydrate and pentahydrate solidify *en bloc* at 30° . The remainder of the curve can be followed in the same manner until the anhydrous ferric chloride is obtained.

It will be observed that if a solution of ferric chloride be evaporated between 30° and 40° , the solution will solidify below 37° , when $\text{Fe}_2\text{Cl}_6 \cdot 12\text{H}_2\text{O}$ is formed as more water is driven off; the mass will liquefy when the eutectic D melting at 27.4° is produced; solidification occurs when the heptahydrate $\text{Fe}_2\text{Cl}_6 \cdot 7\text{H}_2\text{O}$ melting at 32.5° is formed; and the solution will liquefy at 30° when it has the composition corresponding with the eutectic F . The solution will again solidify as the pentahydrate: $\text{Fe}_2\text{Cl}_6 \cdot 5\text{H}_2\text{O}$ melting at 56° appears. Similar phenomena recur with the tetrahydrate, $\text{Fe}_2\text{Cl}_6 \cdot 4\text{H}_2\text{O}$.

If we confine the attention, for a moment, to the curve between the points A and C , it will be apparent that we have the typical curve indicated in Fig. 120, and the same statement applies to the portions CDE , EFG , GHI , and IJK . The point C represents a fused solution of $\text{Fe}_2\text{Cl}_6 \cdot 12\text{H}_2\text{O}$; and the point E a fused solution of $\text{Fe}_2\text{Cl}_6 \cdot 7\text{H}_2\text{O}$. The point D corresponds with the eutectic mixture of both the dodeca- and the heptahydrate. Each maximum point— C , E , G , and I —corresponds with a definite hydrate; and each minimum— B , D , F , H , J —with the eutectic mixture of the two compounds represented by the adjacent maxima. The solubility curve of ferric chloride thus reveals the presence of four stable hydrates, namely $\text{Fe}_2\text{Cl}_6 \cdot 12\text{H}_2\text{O}$; and $\text{Fe}_2\text{Cl}_6 \cdot 7\text{H}_2\text{O}$, $\text{Fe}_2\text{Cl}_6 \cdot 5\text{H}_2\text{O}$, and $\text{Fe}_2\text{Cl}_6 \cdot 4\text{H}_2\text{O}$. Ferric chloride absorbs dry ammonia forming the hexammine, $\text{FeCl}_3 \cdot 6\text{NH}_3$, which becomes the tetrammine, $\text{FeCl}_3 \cdot 4\text{NH}_3$, at 100° , and decomposes at higher temperatures. It is hydrolyzed by water forming ammonium chloride and ferric hydroxide.

Ferrous chloride, FeCl_2 .—Heat metallic iron in a stream of dry hydrogen chloride in a similar apparatus to that employed for the preparation of ferric chloride, Fig. 117. White feathery crystals of anhydrous ferrous chloride are produced. These are very deliquescent. The aqueous solution of ferrous chloride, or the salt formed by dissolving iron, or ferrous carbonate in hydrochloric acid, deposits pale bluish-green crystals of $\text{FeCl}_2 \cdot 4\text{H}_2\text{O}$. Ferrous chloride and ammonia gas at room temperatures form ferrous hexammine chloride, $\text{FeCl}_2 \cdot 6\text{NH}_3$, and at about 300° forms the diammine $\text{FeCl}_2 \cdot 2\text{NH}_3$, which above 350° passes into iron nitride, Fe_3N_2 . When heated in air, ferrous chloride forms ferric oxide and ferric chloride: $12\text{FeCl}_2 + 3\text{O}_2 = 2\text{Fe}_2\text{O}_3 + 8\text{FeCl}_3$. Between 1200° and

1500° in an atmosphere of hydrogen chloride the vapour density corresponds with FeCl_2 , and at lower temperatures possibly with Fe_2Cl_4 .

Aluminium chloride, AlCl_3 .—The aqueous solution, made by dissolving the metal in hydrochloric acid, deposits crystals of $\text{AlCl}_3 \cdot 6\text{H}_2\text{O}$. These crystals decompose when heated, forming alumina, etc.: $2(\text{AlCl}_3 \cdot 6\text{H}_2\text{O}) = 6\text{HCl} + 9\text{H}_2\text{O} + \text{Al}_2\text{O}_3$. The crystalline hydrate cannot be prepared by the evaporation of the aqueous solution unless in a stream of hydrogen chloride, because of hydrolysis (see magnesium chloride). Anhydrous aluminium chloride is made on a large scale by passing chlorine over a mixture of carbon and alumina at a high temperature. The volatile product is condensed in a receiver to which no moisture has access. Neither the chlorine nor the carbon can alone react with the oxide under the conditions of the experiment. The joint action of the carbon and chlorine is needed for the work: $\text{Al}_2\text{O}_3 + 3\text{C} + 3\text{Cl}_2 = 2\text{AlCl}_3 + 3\text{CO}$.

Anhydrous aluminium chloride can also be prepared either by heating aluminium turnings in a current of dry chlorine; or in a rapid stream of dry hydrogen chloride, say, in the apparatus illustrated in Fig. 117. The aluminium turnings should be cleaned free from grease by washing with alcohol and ether, and finally drying at 150°. When all the air has been expelled from the apparatus by the stream of hydrogen chloride, heat the bulb containing the aluminium. The metal reacts with the gas: $2\text{Al} + 6\text{HCl} = 2\text{AlCl}_3 + 3\text{H}_2$. The aluminium chloride passes into the receiver. It may be necessary to heat the tube leading into the receiver from time to time to prevent the blocking of the tube by the condensation of aluminium chloride.

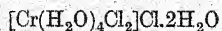
Anhydrous aluminium chloride sublimes without fusion at 183°, forming white hexagonal crystals. If heated under pressure, aluminium chloride melts before volatilization. The anhydrous chloride fumes strongly in moist air: $\text{AlCl}_3 + 3\text{H}_2\text{O} = \text{Al}(\text{OH})_3 + 3\text{HCl}$. The heat of solution of anhydrous chloride is $\text{Al}_2\text{Cl}_6 + \text{Aq} = \text{Al}_2\text{Cl}_{6\text{aq}} + 153.7 \text{ Cals.}$, and the corresponding value for ferric chloride, Fe_2Cl_6 , is + 63.4 Cals. The vapour density at about 200° corresponds with Al_2Cl_6 , and above 450° with AlCl_3 —below this temperature the state of the vapour is represented by $\text{Al}_2\text{Cl}_6 \rightleftharpoons 2\text{AlCl}_3$. Aluminium chloride forms double salts with ammonia, phosphorus pentachloride, alkaline chlorides, etc. The salt $\text{AlCl}_3 \cdot 3\text{NaCl}$ was used by Deville for the preparation of metallic aluminium. An aluminous chloride, AlCl_2 , analogous with ferrous chloride is not known.

Manganous chloride, MnCl_2 .—This salt is prepared by dissolving the oxide or carbonate in hydrochloric acid, and evaporating the solution. Rose-red or pink monoclinic crystals of $\text{MnCl}_2 \cdot 4\text{H}_2\text{O}$ are obtained. The anhydrous salt is made by heating the crystals in a stream of hydrogen chloride. Manganous chloride forms double salts with the alkaline chlorides. The vapour density between 1200° and 1500° corresponds with the formula, MnCl_2 .

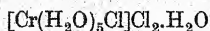
Manganic chloride.—A cold solution of manganese dioxide in concentrated hydrochloric acid behaves as if it contained a higher chloride than MnCl_2 . The solution probably contains a large proportion of MnCl_3 ; some think MnCl_4 is also present. Manganese trichloride, MnCl_3 , is made by treating manganese dioxide, suspended in carbon tetrachloride, with dry hydrogen chloride. The precipitated solid is a mixture of manganese tri- and tetra-chlorides, which is decomposed by water; but when the mass is extracted with anhydrous ether, and the ether evaporated, manganese trichloride remains as a black powder with a greenish tinge.

It is immediately hydrolyzed by water. The residue from the ether extraction is manganese tetrachloride, MnCl_4 , which is soluble in absolute alcohol. The alcoholic solution slowly decomposes. The rapid decomposition of these chlorides in aqueous solution explains the hypothetical character of our statements about the higher chlorides of manganese until a solvent was discovered in which they could be dissolved without rapid decomposition. The isomorphism of the following pairs of salts: $\text{MnCl}_3 \cdot 2\text{NH}_4\text{Cl} \cdot \text{H}_2\text{O}$ with $\text{FeCl}_3 \cdot 2\text{NH}_4\text{Cl} \cdot \text{H}_2\text{O}$; and of $\text{FeCl}_3 \cdot 2\text{KCl} \cdot \text{H}_2\text{O}$ with $\text{MnCl}_3 \cdot 2\text{KCl} \cdot \text{H}_2\text{O}$, indicates a certain analogy between iron and manganese.

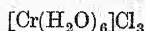
Chromic chloride, CrCl_3 .—The anhydrous salt and the hydrate $\text{CrCl}_3 \cdot 6\text{H}_2\text{O}$ are prepared by methods similar to those employed for aluminium chloride. Free chlorine can be detected in the vapour of chromic chloride at 355° . A number of different hydrates have been made by dissolving chromic hydroxide in hydrochloric acid, or by reducing the chromates and dichromates in solutions of hydrochloric acid. The heat of solution of the green salt which enters into reversible reaction with the violet salt in aqueous solutions is 31.5 Cals., and of the violet salt, 22.2 Cals., so that the heat of transformation from the green to the violet salt is -9.3 Cals. The green solutions pass slowly into the violet solutions. The speed of the conversion has been measured. The reaction is reversible. The violet salt appears to be stable in concentrated solutions, and the green salt in dilute solutions. Crystals of the insoluble violet form of chromic chloride can be left several days under water without appreciable action; but if the merest trace, say, 0.0002 grm. of chromous chloride be present, the temperature of the mixture rises, and the crystals dissolve, forming an indigo-blue solution. This is a curious catalytic reaction. The change is not due to a difference in the degree of hydration of the chromic chloride, or to the formation of a basic salt. Other chromic salts exhibit an analogous change. There are three varieties of the hexahydrated chloride: monoclinic crystals of a violet colour—the β -salt—made by passing hydrogen chloride into a saturated solution of the oxide in hydrochloric acid at a low temperature 8° to 10° ; at a higher temperature, the violet solution becomes green, and when saturated with hydrogen chloride, it deposits rhombic green crystals—the α -salt. Another green salt—the γ -salt—is obtained by treating the mother liquid from the violet crystals with ether saturated with hydrogen chloride. These three salts are remarkable in that when freshly prepared solutions are treated with silver nitrate with the α -salt one-third the total chlorine is precipitated as silver chloride; with the γ -salt two-thirds; and with the violet β -salt all is precipitated. This is symbolized by Werner's formula to be described later:



Green α -salt



Green γ -salt



Violet β -salt.

The violet crystals of the hexahydrate are practically insoluble in water unless a trace of chromous chloride, CrCl_2 , or certain other reducing agents, be present.

Chromous chloride, CrCl_2 .—According to Moberg, when chromic chloride is heated in a current of hydrogen until hydrogen chloride ceases to be evolved, a white crystalline compound, CrCl_2 , is formed. Moissan made it by heating chromic chloride in the vapour of ammonium chloride; or

the metal in hydrogen chloride. The crystals dissolve in water forming a blue solution which rapidly absorbs oxygen and forms the green chromic chloride. A solution of chromous chloride is formed by dissolving the metal in hot hydrochloric acid. The green solution is reduced to the blue solution of chromous chloride by the action of zinc.

Cobaltous chloride, CoCl_2 .—The hydrated salt— $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ —crystallizes from solutions of the oxide or carbonate in hydrochloric acid in the form of ruby-red monoclinic crystals which are easily soluble in water. The crystals at about 100° form the monohydrate, $\text{CoCl}_2 \cdot \text{H}_2\text{O}$; and at about 120° , the anhydrous chloride CoCl_2 . The latter is deep blue but gives reddish solutions when dissolved in water. Cobaltic chloride, CoCl_3 , is known in aqueous solution.

Nickelous chloride, NiCl_2 .—Anhydrous nickel chloride, NiCl_2 , is formed in a similar manner to the cobalt salt. The golden-yellow crystals of anhydrous salt dissolve slowly in water. The aqueous solution deposits green monoclinic prismatic crystals, $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$, which are inclined to effloresce in air. Nickelic chloride, NiCl_3 , is not known.

Questions.

1. Describe the preparation of chlorine and hydrogen chloride, and contrast their properties. What is the relative density of the two substances?—*Aberdeen Univ.*

2. Write a short essay on one of the following subjects: (a) Catalysis and catalytic agents; (b) The chemical work of Priestley and Lavoisier; (c) The determination of the atomic weight of chlorine.—*Aberystwyth Univ.*

3. Mention four elements that will liberate hydrogen from hydrochloric acid, and write equations of reactions.—*Sheffield Scientific School, U.S.A.*

4. State the law of combination of gases by volume and Avogadro's hypothesis. From a consideration of the combination of gaseous hydrogen and chlorine how can it be shown that the molecule of hydrogen consists of parts?—*Univ. North Wales.*

5. What would be the cost of materials in preparing a 20 per cent. solution of hydrochloric acid if 1 kilo of sodium chloride costs 2d., 1 kilo of 90 per cent. sulphuric acid costs 3d., and a kilo of Glauber's salt, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$, can be sold for 1½d. per kilo?—*F. Rudorff, Grundriss der Chemie.*

6. What kinds of salts hydrolyze in water? Give three equations illustrating hydrolysis.—*Sheffield Scientific School, U.S.A.*

7. What is meant by the dissociation of a gas? How is dissociation distinguished from ordinary decomposition? Discuss the dissociation phenomena exhibited by nitrogen peroxide under the influence of heat.—*Sheffield Univ.*

8. Describe the preparation of zinc chloride from zinc oxide. Calculate the volume of a solution of hydrochloric acid, density 1.1 and containing 20 per cent. by weight of acid required to make 50 grams of this salt.—*Sheffield Scientific School, U.S.A.*

9. Describe fully how you would prepare (a) a solution of hydrogen chloride, (b) a solution of chlorine, from common salt. How could it be shown that the solution of the acid is free from chlorine?—*Univ. North Wales.*

10. How did Faraday effect the liquefaction of chlorine? Give a short sketch of recent researches upon the liquefaction of gases.—*London Univ.*

11. What discoveries are associated with the names of Priestley, Davy, Faraday, Graham?—*Aberdeen Univ.*

12. Write the formulae you are accustomed to use for the chlorides of potassium, silver, aluminium, tin, mercury, and iron: which of these expressions may be regarded as representing molecular composition, and upon what grounds may they be so regarded?—*London Univ.*

13. Plot the specific gravities of solutions of hydrochloric acid of different concentration from the following table:—

Grms. HCl per litre	12	96	152	273	303	404	469
Specific gravity	1.005	1.045	1.070	1.125	1.135	1.175	1.200

Show (a) the weight of HCl in a solution of specific gravity 1.180; and (b) the per cent. of hydrochloric acid in a solution of specific gravity 1.200.

14. According to the determinations of Stas, KClO_3 contains 39.15 per cent. of oxygen; and 100 grams of silver are precipitated by 69.1 grams of KCl: 100 grams of silver form 132.85 grms. of AgCl . From these data calculate the atomic weight of chlorine on the assumption that the atomic weight of oxygen is 15.9.—*New Zealand Univ.*

15. Write a short account of the phase rule with special reference to its application to the fusion, efflorescence, and deliquescence of hydrated salts.—*Panjab Univ.*

16. 0.2925 grm. of a specimen of common salt from the Khewrah mines was dissolved in water and completely precipitated with a solution of silver nitrate. The precipitate after washing and drying was found to weigh 0.7075 grm. What was the percentage of sodium chloride in the specimen?—*Panjab Univ.*

CHAPTER XIV

THE RELATIONS OF CHLORINE—IODINE, BROMINE, AND FLUORINE

§ 1. Bromine.

Atomic weight, Br = 79.92; molecular weight, Br₂ = 159.84; valency, 1-, 3-, 5-, 7-valent; critical temperature, 302.2°; melting point, - 7.3°; boiling point, 59°. Specific gravity of the liquid at 0°, 3.1883. Vapour density: (H₂ = 2), 158.85; (air = 1), 5.524.

History.—While studying the mother liquid which remains after the crystallization of salt from the water of the salt marshes of Montpellier, A. J. Balard (1826) was attracted by the intense yellow coloration developed when chlorine water is added to the liquid. Balard digested the yellow liquid with ether; decanted off the supernatant ethereal solution; and treated this with potassium hydroxide. The colour was destroyed. The solution was evaporated to dryness. The residue resembled potassium chloride, but unlike the chlorides, when heated with manganese dioxide and sulphuric acid, it furnished red fumes which condensed to a dark brown liquid with an unpleasant smell.¹ Balard called this substance "muride," but afterwards changed the name to "bromine"—from the Greek *βρῶμος* (bromos), stench. Balard demonstrated the elementary nature of bromine, and showed its relation to iodine and chlorine.

Occurrence.—Bromine does not occur free in nature. Small quantities occur combined in many silver ores; and it occurs associated with potassium, sodium, magnesium, or calcium in many mineral waters, salt springs—e.g. the Congress and Excelsior springs of Saratoga, Woodhall Spa, etc.—and sea-water. The water of the Atlantic is said to contain 0.007 per cent. of magnesium bromide; the water of the Dead Sea, 9 per cent.; the mineral water of Ohio, from 3.4 to 3.9 per cent. of magnesium bromide. The ratio of chlorides to bromides in the marine waters of the globe is almost constant, except for the Black and Baltic Seas. It has also been estimated that there are about 120000,000000 tons of bromides present in the marine waters of the earth. Bromine is also found in marine animals

¹ J. von Liebig used to relate that some years before Balard's discovery he received, from a salt manufactory in Germany, a vessel containing bromine, or at least a product very rich in bromine, with a request to examine it. Believing the liquid to be iodine chloride, he did not subject the specimen to a very exhaustive study. When he heard of the discovery of Balard, Liebig saw his mistake, and placed the vessel in a special cabinet for storing mistakes—*l'armoire des fautes*. Liebig pointed this out to his friends to show how easily one could get very close to a discovery of the first rank and yet fail to grasp the facts when guided by preconceived ideas.

and plants, in rock salt, etc., human urine, and all commercial products directly or indirectly derived from sea-salt or Stassfurt salts. The saline deposits of Stassfurt contain about one per cent. of magnesium bromide. From this latter source, the main supply of the bromine in commerce is derived.

Manufacture.—Perhaps two-thirds of the world's annual consumption of bromine is obtained in Germany from the mother liquid remaining after the separation of the potassium salts, p. 270, which contains about 0.25 per cent. of bromine in the form of magnesium bromide. The hot liquid is allowed to percolate down a tower packed with earthenware balls; here the descending liquid meets an ascending stream of chlorine gas. The magnesium bromide is decomposed: $\text{MgBr}_2 + \text{Cl}_2 = \text{MgCl}_2 + \text{Br}_2$. The bromine vapours leave the top of the tower through an exit pipe, and are liquefied in a suitable condenser. In the electrolytic process, the mother liquid is electrolyzed. The magnesium bromide can be decomposed by the electric current before the magnesium chloride, and if any chlorine were formed it would at once decompose the magnesium bromide as just indicated. It is thus possible and practicable to separate the liberated bromine.

Purification.—The bromine is purified by redistillation. The chlorine is removed by distillation from calcium or ferrous bromide, or by collecting separately in a receiver the first 1 to 4 per cent. of the distillate in the form of a volatile compound or mixture of chlorine and bromine. If anhydrous bromine be needed, it must be redistilled off concentrated sulphuric acid. If iodine be present, this must be removed by treatment with a copper salt. Cuprous iodide, CuI , is precipitated.

Properties.—At ordinary temperatures, bromine is a heavy mobile reddish-brown liquid—specific gravity at 0° , 3.1883. It freezes to a yellowish-brown crystalline mass at -7.3° , and boils in air at 59° . It is very volatile, and gives off a dark reddish-brown vapour at ordinary temperatures. Bromine separates in the form of carmine-red needle-like crystals, when a solution of bromine in carbon disulphide is cooled to -90° . The solid is almost colourless when cooled to -252° . Bromine has a disagreeable irritating smell, and it attacks the eyes as well as the mucous membrane of the throat and nose. It is poisonous, and attacks the skin, producing troublesome sores. 100 grams of water dissolve 43 grams of bromine at 0° , and 32 grams at 20° . Bromine is also soluble in chloroform, carbon disulphide, alcohol, ether, and acetic acid. When bromine water is cooled to 0° , it forms bromine hydrate, $\text{Br}_2 \cdot 10\text{H}_2\text{O}$. Bromine resembles chlorine in general chemical properties, but is not quite so energetic. It forms bromides analogous with chlorides, and only unites slowly with hydrogen in sunlight.

Atomic and molecular weights.—Analyses of silver bromide show that the experimental value of the combining weight of bromine is very close to 79.92 when silver is 107.88. This number, 79.92, also represents the atomic weight of bromine, because it represents the smallest amount of bromine in any one of the known volatile compounds of bromine.

At about 100° the vapour density is 5.87 (air = 1); and at 228° , 5.52. This corresponds with the molecule Br_2 . At 1570° the vapour density is but two-thirds of its value at 228° , showing that the molecule is probably dissociating into atoms.

Uses.—Bromine is used in metallurgy, photography, and the chemical industries. It is also used as a disinfectant, for which purpose it is sold under the name "bromum solidificatum," which is kieselguhr—a siliceous earth—saturated with bromine. Bromine is also used as an oxidizing agent.

§ 2. Hydrogen Bromide—Hydrobromic Acid.

Molecular weight, $\text{HBr} = 80.93$. Melting point, -87.9° ; boiling point, -68.7° ; critical temperature, 91° ; Vapour density ($\text{H}_2 = 2$), 80.23 ; (air = 1) 2.79 . One litre of the gas under normal conditions weighs 3.608 grams; specific gravity of liquid, 1.63 at 10° .

Hydrogen bromide.—This gas can be obtained directly from its elements. If a stream of hydrogen be bubbled through warm bromine, the hydrogen which passes along is highly charged with bromine vapour, and when ignited, dense clouds of hydrogen bromide are formed. If a mixture of

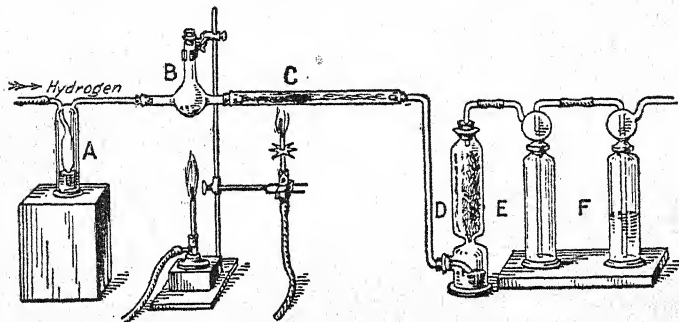


FIG. 121.—Catalytic Process for Hydrogen Bromide.

hydrogen and bromine be passed through a red-hot tube containing platinized asbestos, or a hot platinum spiral, the elements rapidly combine. The platinum acts as a stimulant or catalytic agent. An apparatus for this experiment is illustrated in Fig. 121. The hydrogen is dried by passage through concentrated sulphuric acid in the wash-bottle *A*; and bromine is placed in the bulb *B*. When all the air is expelled from the apparatus, the bromine is warmed slightly, and the mixed gases are passed through the tube *C* containing platinized asbestos. Any bromine which escapes uncombined is absorbed by red phosphorus, slightly damped, which is packed along with glass wool in the tower *D*; here any free bromine forms a phosphorus bromide. If the hydrogen bromide is to be absorbed by water, the products of the reaction—a mixture of hydrogen and hydrogen bromide—are passed through an empty wash-bottle *E*, and then through a similar wash-bottle half filled with water. The empty wash-bottle *E* is reversed so that there is no danger of water passing beyond the bottle if back suction occurs. Hydrogen bromide is also formed when some of the metallic bromides are reduced to metals in a current of hydrogen. Thus, at a red heat with silver bromide: $2\text{AgBr} + \text{H}_2 = 2\text{HBr} + 2\text{Ag}$. Hydrogen bromide is sometimes made by the action of bromine on hydrocarbons—*e.g.* naphthalene, C_{10}H_8 ; benzene, C_6H_6 ; anthracene, $\text{C}_{10}\text{H}_{10}$;

etc.—but the gas is then more or less contaminated with organic products. The gas is conveniently made by the action of bromine on phosphorus and water.

The following is the method usually employed in the laboratory: Mix, say, 10 grams of red phosphorus with 80 grams of fine sand, and place the dry mixture in a dry distillation flask, *A*, Fig. 122. Add about 20 c.c. of water. Close the flask with a rubber stopper fitted with a tap funnel *B*, and delivery tube as indicated in Fig. 121. The end of the tap funnel is drawn to a fine point. Connect the delivery tube with a tower or a U-tube *C* containing glass wool and slightly damp red phosphorus. The gas may be collected by the upward displacement of air, or absorbed in water as illustrated in the previous diagram; or collected over mercury, *E*. In the latter case, a safety funnel may be attached to the delivery tube at *D*, so that variations of pressure inside the apparatus may be rapidly adjusted without risk of explosion and back suction of mercury. About 60 grams of bromine are placed in the tap funnel, and allowed to fall, drop by drop, on the red phosphorus. As each drop of bromine comes in contact with the phosphorus,

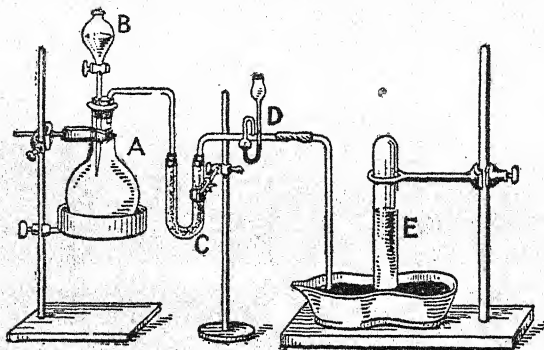


FIG. 122.—Preparation of Hydrogen Bromide.

gently warming the flask. The hydrogen bromide can be dried by means of a tube packed with calcium bromide.

The chemical reactions which occur during the preparation of hydrogen bromide by the action of bromine on phosphorus and water, are probably somewhat as follows: Phosphorus tri- and penta-bromide are first formed; these react with the water: $\text{PBr}_3 + 3\text{H}_2\text{O} = \text{H}_3\text{PO}_3 + 3\text{HBr}$; and $\text{PBr}_4 + 4\text{H}_2\text{O} = \text{H}_3\text{PO}_4 + 5\text{HBr}$. The whole reaction is usually represented on the supposition that phosphorus tri-bromide is formed: $2\text{P} + 6\text{H}_2\text{O} + 3\text{Br}_2 = 2\text{H}_3\text{PO}_3 + 6\text{HBr}$. If too little water be present, some crystals of phosphonium bromide— PH_4Br —may be formed in the flask owing to the decomposition of the hot phosphorous acid, H_3PO_3 : thus: $4\text{H}_3\text{PO}_3 = 3\text{H}_3\text{PO}_4 + \text{PH}_3$; and $\text{PH}_3 + \text{HBr} = \text{PH}_4\text{Br}$.

Hydrobromic acid.—It is convenient to call the gas “hydrogen bromide,” and aqueous solutions “hydrobromic acid.” Hydrobromic acid is formed when hydrogen sulphide is passed into bromine covered with a layer of water: $\text{H}_2\text{S} + \text{Br}_2 = 2\text{HBr} + \text{S}$. If sulphur dioxide be used in place of hydrogen sulphide a pale yellow homogeneous liquid is obtained: $\text{SO}_2 + \text{Br}_2 + 2\text{H}_2\text{O} \rightleftharpoons 2\text{HBr} + \text{H}_2\text{SO}_4$. When this liquid is distilled in, say, the apparatus depicted in Fig. 123, an aqueous solution of hydrogen bromide is obtained. The reaction is reversible sulphuric

a flash of light is produced. Some prefer to keep the flask immersed in cold water during the early stages of the reaction, and to wrap a towel round the flask in case of an explosion. The heat of the reaction volatilizes some bromine which is retained by the phosphorus in the U-tube; when all the bromine has been run into the flask, a further quantity of hydrogen bromide may be obtained by

acid is reduced by hydrogen bromide forming bromine and sulphur dioxide. This helps us to understand why hydrogen bromide cannot be satisfactorily prepared by the action of sulphuric acid upon potassium bromide as in the preparation of hydrogen chloride by the action of sulphuric acid upon potassium chloride. When the attempt is made, colourless hydrogen bromide is first given off, but the issuing gas immediately acquires a yellow colour, and then a brown tinge, showing that bromine is also evolved. The issuing gas also contains sulphur dioxide. Hence not only does the reaction $KBr + H_2SO_4 = KHSO_4 + HBr$ take place, but also the consecutive reaction $2HBr + H_2SO_4 = SO_2 + Br_2 + 2H_2O$, and the method is accordingly impracticable. If dilute sulphuric acid be employed with the idea of preventing an appreciable decomposition of the hydrogen bromide, the amount of hydrogen bromide obtained is very small. If phosphoric acid be used in place of sulphuric acid because phosphoric acid does not de-oxidize so readily as sulphuric acid, the action— $KBr + H_3PO_4 = KH_2PO_4 + HBr$ —is rather slow.

Properties.—Hydrogen bromide is a colourless gas with a strong penetrating smell. It fumes in air. The gas condenses to a liquid at -73° ; and solidifies to a colourless solid at -87° . The gas is very soluble in water. A solution saturated at 0° has a specific gravity 1.78, and the hydrogen bromide and water are approximately in the proportion $HBr : H_2O$. 100 grams of water at 0° dissolve 221 grams of hydrogen bromide, and 199 grams at 20° . The acid containing 48 per cent. of hydrogen bromide distills unchanged in composition; weaker acids lose much water until the constant boiling acid is obtained, which distills at 126° and 760 mm. pressure; more concentrated acids lose hydrogen bromide until the constant boiling acid is obtained. Two crystalline hydrates— $HBr \cdot H_2O$ and $HBr \cdot 2H_2O$ —have been made.

Dry hydrogen bromide has no action on litmus. Hydrogen bromide resembles hydrogen chloride very closely in chemical properties, but hydrogen bromide is less energetic. It attacks mercury very slowly, forming hydrogen and mercury bromide. The action is not fast enough to interfere with the experiment, Fig. 122. Hydrobromic acid forms bromides in the same way that hydrochloric acid forms chlorides. Hydrogen bromide is more easily decomposed than hydrogen chloride. At 800° decomposition is quite appreciable. The gas is also slightly decomposed by exposure to sunlight. The composition of hydrogen bromide has been determined in a similar manner to that of hydrogen chloride.

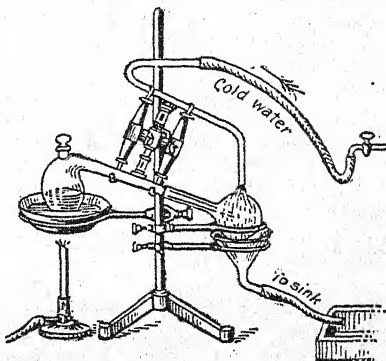


FIG. 123.—Distillation of Hydrobromic Acid.

§ 3. Iodine.

Atomic weight, $I = 126.92$; molecular weight, $I_2 = 253.84$; valency 1-, 3-, 5-, 7-valent. Melting point, 113° ; boiling point, 184.35° ; critical temperature, 553° . Specific gravity of the solid, 4.933. Vapour density, 253.84 ($H_2 = 2$), and 8.72 (air = 1).

History.—In 1812, B. Courtois, a manufacturer of saltpetre, near Paris, used an aqueous extract of varec or kelp¹ for decomposing the calcium nitrate from the nitre beds, as indicated later on. Courtois noticed that the copper vats in which the nitrate was decomposed were rapidly corroded by the liquid, and he traced the effect to a reaction between the copper and an unknown substance in the lye obtained by extracting the varec or kelp with water. On evaporating the aqueous extract of the kelp, crystals of potassium sulphate first separate, then follow crystals of sodium sulphate, sodium chloride, and afterwards sodium carbonate. The remaining liquid when heated with sulphuric acid in a retort furnished “a vapour of a superb violet colour” which condensed in the beak of the retort and in the receiver in the form of brilliant crystalline plates. Courtois communicated his discovery to Clément and Désormes, who published some results of their study of this new substance in 1813. A year later Gay-Lussac published an extensive and remarkable memoir on this new substance which was called iodine—from the Greek *ioeîdēs* (ioeidēs), violet. Gay-Lussac established the elementary nature of iodine, and demonstrated its relationship to chlorine. About the same time, H. Davy confirmed many of Gay-Lussac’s results.

Occurrence.—Varec or kelp contains from 0.1 to 0.3 per cent. of iodine. The kelp derived from deep seas is richer in iodine than that from shallower parts. Iodine also occurs in small quantities in sea water; sea plants; sea animals; in some land plants and animals; in codliver oil; in the thyroid gland of animals; and in many mineral springs. It occurs combined with silver in some Mexican ores, and in some South American lead ores. Most of the iodine of commerce is extracted from the mother liquid remaining after the separation of sodium nitrate from caliche in Peru, etc. Caliche contains about 0.2 per cent. of iodine, and the mother liquid after the extraction of the sodium nitrate, contains 5 to 20 per cent. of sodium iodate.

Manufacture.—The mother liquid remaining after the crystallization of the sodium nitrate from the aqueous extract of caliche is treated with sodium hydrosulphite, which first reduces the sodium iodate to sodium iodide, and finally to free iodine: $2NaIO_3 + 5NaHSO_3 = 3NaHSO_4 + 2Na_2SO_4 + H_2O + I_2$. The solid iodine which separates is allowed to settle, washed, and pressed into blocks. The impure iodine so obtained is sublimed in iron retorts and the vapours condensed in a series of earthenware receivers.

Iodine is extracted from the ash of seaweed by concentrating the aqueous extract so as to remove alkaline carbonates, chlorides, and sulphates by crystallization. The mother liquid containing the iodides

¹ During the stormy months of spring, seaweeds are washed on to the western coasts of Ireland, Scotland, and France. The inhabitants collect the weed and burn it in large heaps at as low a temperature as possible. The ash thus obtained is called *kelp* in Scotland and *varec* in Normandy.

and some bromides is treated first with sulphuric or hydrochloric acid, and then with manganese dioxide. On heating the mixture, iodine is liberated: $2\text{NaI} + \text{Cl}_2 = 2\text{NaCl} + \text{I}_2$. This is condensed in earthenware receivers arranged so that any water distilled over is condensed and drained off. A ton of kelp is said to furnish 10 to 12 lbs. of iodine.

Purification.—Commercial iodine always contains both chlorine and bromine in solid solution. To purify the iodine, dissolve resublimed iodine in a concentrated solution of potassium iodide; and precipitate the iodine from the solution by adding an excess of water. Wash and dry the solid. The dry precipitate is then intimately mixed with potassium iodide and heated in, say, a beaker covered with a glass flask kept cold by a current of water. A crust of the purified iodine condenses on the bottom, outside the cooled flask. Pure iodine has also been obtained by heating cuprous iodide to about 240° in a current of dry air.

Properties.—At ordinary temperatures, iodine is a dark bluish-black crystalline solid. The rhombic crystalline plates have a metallic lustre, and a specific gravity of 4.933 at 4° . Solid iodine has a vapour pressure of 0.030 mm. at 0° , rising to 3.084 mm. at 55° ; this agrees with the fact that iodine vaporizes slowly at ordinary temperatures, and it has a slight smell resembling chlorine. It is very sparingly soluble in water: 100 c.c. of a saturated solution at 25° contain about 0.034 gram of iodine, and is coloured a faint brown. The iodine is much more soluble if potassium iodide be also present. Iodine is fairly soluble in many organic solvents—alcohol, ether, acetone, chloroform, carbon disulphide, benzene, etc. The colours of the solutions differ with different solvents: hydrocarbons, chloroform, and carbon disulphide give violet solutions; while alcohol, water, aqueous solutions of potassium iodide, and ether give yellow, red, or brown solutions, according to the nature of the solvent and concentration of the solution. The violet solutions are usually obtained with saturated compounds, while brown solutions are obtained with associated liquids having residual valencies—oxygen quadrivalent. This is illustrated by the following experiment: Place a layer of carbon disulphide, *A*, at the bottom of a glass cylinder, Fig. 124; on this a layer of water, *B*, and above all, a layer of ether, *C*. Drop some crystals of iodine into the cylinder. The lowest layer will form a violet solution; water will give a yellowish-brown solution, and the ether a brown solution. Many of the violet-coloured solutions become brown when cooled to low temperatures; and conversely, many brown solutions become violet when heated. It is probable that free iodine is present in the violet solutions; and that the brown coloration is due to the formation of a compound of iodine with the solvent. In brown solutions of iodine in potassium iodide, for example, the iodine is supposed to form an unstable potassium tri-iodide, KI_3 — $\text{KI} + \text{I}_2 \rightleftharpoons \text{KI}_3$. This salt has indeed been isolated in the form of dark-coloured needle-shaped crystals. It is probable that in this salt the potassium is univalent, and the iodine uni- and tervalent, $\text{K} - \text{I} = \text{I}_2$. In the salts RbICl_4 , CsI_5 , etc., the iodine appears to be quinquevalent. A great number of polyiodides and polybromides of the univalent cesium and rubidium have been prepared; they crystallize well, and are more



FIG. 124.—
Colours of
Iodine Solu-
tions.

stable than the corresponding potassium salts: *e.g.* CsI_3 , CsBr_5 , CsI_5 , RbI_3 , RbBr_3 , RbICl_4 , CsI_3 , NH_4I_3 , etc. The tendency of the alkali metals to form higher polyiodides increases with increasing atomic weight in passing from lithium to caesium. Tetraiodides of the alkaline earths have been reported, BaI_4 , SrI_4 , and CaI_4 .

Iodine resembles chlorine and bromine in its chemical properties, but it is rather less energetic. Chlorine can displace bromine from bromides, and both chlorine and bromine can displace iodine from iodides. Chlorine can displace bromine from bromates, and iodine can displace chlorine from chlorates: $2\text{KClO}_3 + \text{I}_2 = 2\text{KIO}_3 + \text{Cl}_2$. Iodine, like chlorine and bromine, combines with many elements, forming iodides. When phosphorus and iodine are placed in contact, the phosphorus melts and inflames, forming either phosphorus triiodide, PI_3 , or phosphorus pentaiodide, PI_5 . Antimony powder inflames spontaneously when it is shaken with iodine vapour; iodine and mercury also combine energetically when heated.

When in contact with starch, iodine forms an intense blue coloration. The reaction is delicate enough to reveal the presence of 0.000001 grm. of iodine per c.c. The blue colour disappears when heated to about 80° , but returns on cooling. The "blued" starch is supposed to be either a solid solution of iodine in the starch, or else an "addition" compound of iodine with starch.

Uses.—Iodine is used in medicine, the manufacture of dyes and organic compounds, in photography, and in analytical chemistry.

Atomic and molecular weight.—The combining weight of iodine, determined from the analysis of silver iodide and other iodine compounds, ranges from 126.79 to 126.93, and the best representative value is supposed to be 126.92 when oxygen is 16. This is the smallest amount of iodine in all known volatile compounds containing iodine, and hence this number is taken to be the atomic weight of iodine. This agrees with a two-atom molecule, for the vapour density of iodine, at 600° , is 8.7 (air = 1); or 125.9 ($\text{H} = 1$). The lowering of the freezing point and the raising of the boiling point of solutions of iodine in carbon disulphide acetic acid and chloroform show that the molecular weight is nearly 254, corresponding with I_2 .

When iodine vapour is heated above 700° its density diminishes steadily up to about 1700° , when it becomes constant at half its value at the lower temperature. Thus the theoretical density for I_2 is 8.78, and for I , 4.39.

Temperature	480°	850°	1043°	1275°	1390°	1468°
Vapour density	8.74	8.07	7.01	5.82	5.27	5.06
Dissociation	0	8.6	25.0	50.5	66.2	73.1 per cent.

Without doubt, the iodine molecule, I_2 , dissociates into atoms: $\text{I}_2 = \text{I} + \text{I}$. Related phenomena will be discussed later.

§ 4. Equilibrium, and the Kinetic Theory of Chemical Action.

" . . . The streaming atoms
Fly on to clash together again, and make
Another and another state of things
For ever . . . "

Rejected Addresses.

When a mixture of equal volumes of iodine and hydrogen gases is passed through a red-hot tube, or, better, over finely divided platinum, or

platinized asbestos, Fig. 121, or charcoal, some hydrogen iodide, HI, is formed. If hydrogen iodide gas be treated in a similar way, some iodine and hydrogen are produced. In either case, if the temperature of the tube be 440° , we have approximately 80 per cent. of hydrogen iodide, and 20 per cent. of a mixture of equal volumes of iodine and hydrogen. The only apparent effect of the catalytic agent—platinized asbestos, etc.—is to accelerate the reaction, and if these agents be absent, the time required to make 80 per cent. of hydrogen iodide from the mixture of hydrogen and iodine is much longer. Once this proportion of hydrogen iodide has been formed, the composition of the exit gases remains unchanged, however long the mixture may be heated at 440° , with or without the catalytic agents.

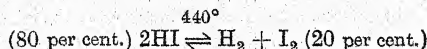
Bimolecular reactions.—We may now extend our previous study of opposing reactions. In the bimolecular reaction



let C_A and C_B respectively denote the concentrations of the substances A and B, expressed in gram-molecules per litre. Similarly, let C_M and C_N respectively denote the concentrations of M and N. We have previously found that the speed of the reaction is equal to the product of the affinity or force driving the reaction, k , and the concentrations of the reacting substances, that is, the velocity of the reaction $A + B$ is equal to $kC_A C_B$. If A and B are the same, so that we have $2A \rightleftharpoons M + N$, the speed of the \rightarrow reaction at any instant will be represented by kC_A^2 . When hydrogen iodide dissociates, we have $2HI \rightleftharpoons H_2 + I_2$. The speed of the \rightarrow reaction at any instant will be represented by kC_{HI}^2 ; and the speed of the \leftarrow reaction by $k'C_I C_I$. When equilibrium occurs, the speeds of these two reactions are the same, and therefore we have the condition of equilibrium:

$$kC_{HI}^2 = k'C_I C_I \text{ or } \frac{k}{k'} = \frac{C_I C_I}{C_{HI}^2} = K$$

At 440° , when the system is in equilibrium, nearly 20 per cent. of the hydrogen iodide will have dissociated. Hence



This means that if 100 molecules of hydrogen iodide be heated to 440° in a closed tube, 20 will have dissociated when the system is in equilibrium. Hence $C_{HI} = 80$; $C_{H_2} = 6400$; and $C_I = C_H = 10$. Hence

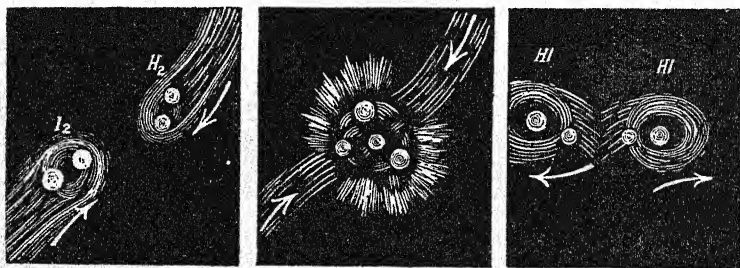
$$\frac{k}{k'} = \frac{1}{64} \text{ or } C_{HI}^2 = 64C_I C_I$$

This means that at 440° , when the concentration of the hydrogen and iodine is unity, these gases will combine 64 times as fast as hydrogen iodide of unit concentration will dissociate, p. 119. It will be observed that each of the direct and reverse reactions is a bimolecular reaction because two molecules are involved in each reaction.

The kinetic theory of chemical action.—The kinetic theory gives an interesting view of chemical action. Imagine a vessel filled with a mixture of equal volumes of iodine and hydrogen gases. The molecules of hydrogen and iodine must be continually crashing together. A *certain proportion* of these collisions will result in chemical change. In the earlier stages of

the reaction, the number of collisions per second between the hydrogen and iodine molecules will be relatively great, but later, as the hydrogen iodide accumulates, the number of collisions between the hydrogen and iodine molecules will become fewer and fewer, and accordingly, the speed of formation of hydrogen iodide will become less and less. Following a plan first used by G. Martin, Fig. 125 may be employed to illustrate a chemically fruitful collision between a hydrogen and an iodine molecule.

Similarly, when two hydrogen iodide molecules crash together, a *certain proportion* of the collisions will result in a dissociation, so that an iodine and a hydrogen molecule will result. At first, the number of collisions will be few and far between, but, as hydrogen iodide accumulates in the system, the number of collisions between these molecules will increase. Finally, when the number of chemically fruitful collisions per second between the iodine and hydrogen molecules is equal to the number of chemically fruitful collisions between the hydrogen iodide molecules, the system will subsequently undergo no perceptible change. Obviously, this does not mean that chemical action has ceased. Every time the



Molecules approach. Molecules collide and react. New molecules separate.

FIG. 125.—Imaginary Representations of the Reaction: $I_2 + H_2 = 2HI$ according to the Kinetic Theory.

proper molecules collide under the right conditions, hydrogen iodide will be formed or dissociated. Both changes proceed with the same velocity, and consequently the composition of the gas as a whole does not alter. Although it is often said that when two opposing reactions are in equilibrium, the system is at a stand-still, the apparent equilibrium is not a state of real tranquillity and repose. Equilibrium is dynamic and active, not static and passive. It is possible to say that a chemical system is in equilibrium only when its composition does not change with the lapse of time.

Dissociation and combination are partial and incomplete.—As the temperature rises the average velocity of the molecules of a gas becomes greater and greater. Although the *average* velocity of the molecules of a gas is constant at any temperature, the velocities of *individual* molecules must vary considerably because of collisions, etc. It is possible that collisions between the *faster moving molecules* of hydrogen iodide, alone, result in dissociation: $2HI = I_2 + H_2$; and that collisions between the slower moving molecules of hydrogen iodide do not produce dissociation; similarly, it may be that collisions between the *slower moving molecules* of

hydrogen and iodine alone result in the formation of hydrogen iodide; and collisions between the fastest molecules do not lead to chemical action. Hence we can see how but a "certain proportion" of the collisions are chemically fruitful. This view of the reaction leads to several other interesting inferences, but since direct proof of the fundamental hypothesis is wanting, sufficient has been given to indicate the trend of modern thought.

The action of stimulants—contact action.—The speed of dissociation of hydrogen iodide at 518° is augmented threefold by raising the pressure from 0.5 to 2 atmospheres. The tremendous condensation of gases on the surfaces of such substances as platinized asbestos, platinum black, charcoal, etc., shows that gases near the surfaces of these substances must be very very concentrated. E. Mitscherlich (1843) estimated that gaseous carbon dioxide condensed on wood charcoal in layers about 0.005 mm. thick, and the gaseous layer is nearly as dense as liquid carbon dioxide. Hence it follows that the concentration of the molecules of one or both the reacting gases must be very great near the surface of the catalytic agent; and, since the temperature is not altered, the average speed of motion will be the same; accordingly, the total number of collisions, and the number of chemically fruitful collisions in unit time will be augmented. This means that the speeds of formation and dissociation of hydrogen iodide will be stimulated in the presence of such substances as platinized asbestos, etc. In other words, these substances act as catalytic agents.

A catalytic agent can alter the speed of a chemical action, but it cannot alter the condition of equilibrium.—Although the speed of a chemical reaction is modified by the presence of a catalytic agent, the final state of equilibrium is not affected. If otherwise, we could allow these substances to react alternately with and without the catalytic agent; this would involve a change in the quantity combined, and the energy thus obtained could be made to do work. This would lead to perpetual motion, which is assumed to be impossible, p. 136. This deduction has been confirmed experimentally with hydrogen iodide with and without platinum black.

Unimolecular reactions.—The last example is instructive. The state of the system in equilibrium will be represented by $kC_{I_2} = k'C_I^2$. If x denotes the proportion of iodine dissociated, and v the volume of the iodine vapour, then, since v volumes of iodine vapour becomes $2v$ volumes of dissociated iodine vapour, it follows that the concentration of the dissociated iodine will be x/v , and of the undissociated iodine $(1-x)/v$. Hence for equilibrium

$$k \frac{1-x}{v} = k' \left(\frac{x}{v} \right)^2 \text{ or } K = \frac{k}{k'} = \frac{x^2}{(1-x)v}$$

In every gram-molecule of iodine (I_2) at 1043° , 0.25 gram-molecule will be dissociated; hence, $x^2 = 0.0625$; $1-x = 0.75$; and $K = 0.0833/v$. To evaluate v , remember that one gram-molecule of iodine vapour at 0° and 760 mm. occupies 22.3 litres; and at 1043° , 107.5 litres. This quantity of gas contains 0.25 more molecules of iodine because of dissociation, and hence its volume is $107.5 + \frac{1}{4}$ of 107.5 = 134.4 litres. Hence $K = 0.0833 \div 134.4 = 0.00062$; or $k : k' = 0.00062 : 1$; or 1 : 1600 (nearly). Otherwise expressed, $C_{I_2} = 1600 C_I^2$, that is, the atoms of iodine will unite 1600

times as fast as the molecules dissociate under such conditions that unit concentration of each is present, p. 315. The dissociation of iodine molecules is a unimolecular reaction because one molecule is concerned in the reaction; and the formation of the two-atom molecule by the union of two one-atom molecules is a bimolecular reaction because two molecules are concerned in the process.

If all the molecules in a reacting system are alike, why do they not all undergo change at the same instant? What regulates the speed of the reaction in such a way that only a certain fraction of the total number of molecules changes in a given time? At first sight, it seems that if all the molecules are in the same condition—either no chemical action will occur, or all the molecules must undergo transformation at the same instant. The answer returned by the kinetic theory is somewhat as follows:

We have just seen that according to the kinetic theory, the average velocity of the molecules becomes greater and greater as the temperature rises; and that although the *average* velocity is constant at any particular temperature, the velocities of *individual* molecules must vary considerably because of collisions, etc. The velocities of the faster moving molecules may finally become so great that the crash, on collision, displaces the atoms from their position of equilibrium in the molecules. Take the case of water, H_2O . Rise of temperature not only accelerates the movements of translation of the molecules themselves, but it probably also increases the amplitudes and velocities of the cyclic motions of the atoms within the molecule. The atomic movements may become so violent that the atoms of one or both molecules are thrown out of the sphere of one another's attraction, and it is no longer possible for the hydrogen and oxygen atoms in the swifter molecules to remain combined. In other words, the molecules may be dissociated. The dissociation of the whole gas is only partial, because the faster moving molecules break down first. There is a kind of struggle for existence, the slower moving molecules are the fittest to survive. It is, however, still possible for the atoms of hydrogen and oxygen to unite each to each, forming elementary molecules; but if a high enough temperature be applied, even the combinations become unstable and two-atom molecules dissociate into one-atom molecules, for the atoms themselves are then free rovers and mutually independent. When the one-atom molecules collide, they enter the sphere of one another's attraction, and, if the velocities of the colliding molecules be not too great, the atoms may remain in contact reforming a two-atom molecule. As before, when the speeds of dissociation and re-combination are equal, the system is in equilibrium.

The relative frequency of uni- and bi-molecular reactions.—Uni- and bi-molecular reactions are very much more frequent than more complex reactions involving three or more molecules. The number of *binary* collisions per second must be very much greater than the number of simultaneous collisions between, say, *three* molecules. When several molecules are involved in a reaction, the reaction must therefore be (1) extremely slow, for they are dependent on the simultaneous juxtaposition of the reacting molecules; or (2) one or more intermediate reactions are involved—one molecule, for instance, may unite with another molecule and the pair (intermediate compound) later collides and reacts with a third

molecule (see "consecutive reactions"); or (3) the reaction takes place in the vicinity of a boundary layer where the reacting molecules form a dense layer and are comparatively close together (see "contact action," p. 317).

§ 5. The Effect of Temperature on Equilibrium—The Principle of Reversibility.

All natural changes take place in such a way that the existing state of things suffers the least possible change.—P. L. M. DE MAUPERTIUS (1747).

Reactions generally proceed faster at a higher temperature than at a lower one, but there are some reactions which are slower at the higher temperature; and at still higher temperatures the action is reversed. The proportion of hydrogen iodide dissociated *decreases* with rise of temperature so long as the temperature does not exceed about 320° ; above that critical temperature, the higher the temperature, the *greater* the amount of hydrogen iodide dissociated. This is illustrated by the graph, Fig. 126. The thermal value of the reaction changes sign at about the same critical temperature; for instance, at 18° , the union of hydrogen and iodine is an *endothermal* reaction: $\text{H}_2 + \text{I}_2 = 2\text{HI} - 6.1$ Cals.; and at 520° , *exothermal*: $\text{H}_2 + \text{I}_2 = 2\text{HI} + 4.4$ Cals. Experience shows that a rise of temperature always favours endothermal reactions, and opposes exothermal reactions. *When a system is in physical or chemical equilibrium, a rise of temperature promotes the formation of those products which are formed with an absorption of heat; a rise of temperature resists the formation of those products formed with an evolution of heat; and a change of temperature has no effect on the equilibrium of reactions thermally neutral.*—J. H. van't Hoff's equilibrium law (1884).

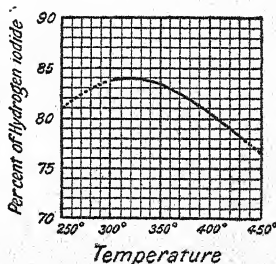


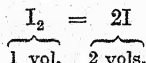
FIG. 126.—Effect of Heat on the Dissociation of Hydrogen Iodide.

Some examples of this law have been previously indicated. The law is simply a special case of the great principle of reversibility. If an exothermal reaction becomes endothermal at a high temperature, we have the curious paradox discussed previously: A compound may be stable at temperatures exceeding that at which it dissociates. The case of hydrogen iodide is particularly instructive. The change in the thermal value of the reaction corresponds with a change in the effect of a rise of temperature on the equilibrium. The principle applies to physical equilibria. When anhydrous sodium sulphate is dissolved in water, heat is evolved, and its solubility is diminished with a rise of temperatures; hydrated sodium sulphate dissolves in water with an absorption of heat, and its solubility increases with rise of temperature. The vaporization of water is an endothermal reaction, and hence a rise of temperature favours vaporization, for it increases the concentration of the vapour phase.

The effect of pressure on equilibria.—The principle is also applicable with other forms of energy. Thus: *when a system is in a state of physical*

or chemical equilibrium, an increase of pressure favours the system formed with a decrease in volume; a reduction of pressure favours the system formed with an increase in volume; and a change of pressure has no effect on a system formed without a change in volume—G. Robin's law (1879). Thus hydrogen iodide is formed from hydrogen and iodine without a change in volume, and the state of equilibrium is not affected by variations of pressure. When ice melts, the liquid occupies a smaller volume than an equivalent amount of ice; experiment shows that the melting point of ice is lowered by pressure in agreement with the law. With sulphur the converse is true. The melting point of sulphur is raised by pressure, but the liquid phase has a greater specific volume than the solid phase.

The principle of least action, underlying these rules, is of wide application and of great service. It gives no answer to the question: What is the magnitude of the change? In consequence, like the phase rule, the principle is qualitative, not quantitative. It virtually says that natural changes take place in such a way that the existing state of things suffers the least possible change; or, as W. D. Bancroft expressed it, a system tends to change so as to minimize the effects of an external disturbance. This has been called "the principle of the opposition of reactions to further change," or, *if a system in physical or chemical equilibrium be subjected to a stress involving a change of temperature, pressure, concentration, etc., the state of the system will automatically tend to alter so as to undo the effect of the stress*—H. le Chatelier's law (1888). For instance, if the temperature of a system in equilibrium be raised a few degrees, the state of the system will change so as to induce the formation of that component or phase which absorbs most heat, and accordingly tend to lower the temperature. If the \rightarrow reaction be exothermal, the change will proceed in the reverse direction; and if the \rightarrow reaction be endothermal, the system will change in the same direction. Again, if the pressure of the dissociating iodine



be increased, the state of the system will change so that the volume is diminished; and conversely, if the pressure be reduced, the state of the system will change so that the volume is increased, that is, the less the pressure the greater the amount of iodine dissociated. In the case of solutions, an increase of concentration will induce the formation of that component or phase which will lower the concentration of the solute added; and an increase of vapour pressure will lead to the formation of that component or phase which will reduce the vapour pressure, etc.

§ 6. Hydrogen Iodide—Hydriodic Acid.

Molecular weight, HI = 127.93. Melting point, -50.8° ; boiling point, -34.1° ; critical temperature, 151° . Vapour density ($H_2 = 2$), 127.67; (air = 1), 4.44. Specific gravity of liquid, 2.27 at 12° .

Hydrogen iodide.—As in the case of the corresponding bromine compound, hydrogen iodide can be made by the direct combination of iodine and hydrogen. It can be made by the action of iodine, on certain organic

compounds—colophonium (resin), copaiva oil, etc. It is also formed by the action of hydrogen on silver iodide. As with hydrogen bromide, but not with hydrogen chloride, hydrogen iodide cannot be made satisfactorily by the action of sulphuric acid upon the corresponding potassium salt—in the present case, potassium iodide. Sulphuric acid, according to the conditions of the experiment, can be reduced by hydrogen iodide to sulphur dioxide, free sulphur, and to hydrogen sulphide. In the latter case, $\text{H}_2\text{SO}_4 + 8\text{HI} \rightleftharpoons 4\text{H}_2\text{O} + 4\text{I}_2 + \text{H}_2\text{S}$. The reaction is reversible so that by passing hydrogen sulphide into iodine-water, a dilute solution of hydrogen iodide can be obtained. Phosphoric acid, in place of sulphuric acid, gives very fair results with potassium iodide. Gaseous hydrogen iodide is usually made by the decomposition of phosphorus iodide by the action of water. This is done by mixing red phosphorus and iodine in a dry flask, and gradually adding water from a dropping funnel to the products of the reaction. This is a modification of the process employed for the preparation of hydrogen bromide, rendered necessary because bromine is liquid, and iodine solid. Free iodine is removed from the gas by passing the hydrogen iodide through a tower packed with red phosphorus and glass-wool. A little phosphonium iodide may be formed in the tower. The gas can be dried by passing it through a tube containing calcium iodide or phosphorus pentoxide. Hydrogen iodide cannot be collected over mercury because the mercury is attacked; it is usually collected by the upward displacement of air.

Hydriodic acid.—The term “hydrogen iodide” is reserved for the gas, and hydriodic acid for the aqueous solution. An aqueous solution of hydrogen iodide can be made by the following process:

Add about 3 grams of powdered iodine to 250 c.c. of water in a 500 c.c. flask and pass a stream of hydrogen sulphide slowly into the mixture. In a few minutes all the iodine will have dissolved owing to the reaction: $\text{H}_2\text{S} + \text{I}_2 = \text{S} + 2\text{HI}$. Add more powdered iodine, and continue the passage of the gas. Repeat the operations until about 20 grams of iodine have been added. Transfer 30 grams more iodine—50 grams in all—to the flask. In about half an hour the iodine will all have dissolved in the hydrogen iodide already formed. Continue passing hydrogen sulphide until the brown colour of the solution disappears, showing that all the iodine has been transformed into hydrogen iodide. Pass a rapid stream of carbon dioxide or hydrogen through the warm solution to drive off the hydrogen sulphide. Shake the solution to coagulate the sulphur; and remove the sulphur from the solution by filtration through glass-wool. The solution can be further purified by distillation; collect the fraction which boils between 125° and 130° . This solution contains about 50 per cent. of hydrogen iodide. A more concentrated solution can be made by passing gaseous hydrogen iodide into cold water, or, better, into a solution of hydrogen iodide made as just described.

Properties.—Hydrogen iodide is a colourless gas which fumes strongly in air. It condenses to a colourless liquid at 0° under 4 atmospheres pressure. This boils at -34.14° , and freezes to a white solid which melts at -50.8° . The gas is very soluble in water: one volume of water at 10° dissolves about 425 volumes of hydrogen iodide. The solution fumes strongly in air, and it has acid properties. The aqueous solution, containing 57 per cent. of hydrogen iodide, boils at 127° (774 mm.), and distils unchanged in composition; weaker acids become stronger, and stronger acids become weaker on boiling until the 57 per cent. acid is obtained, when the solution distils unchanged in composition. Hydriodic acid is colourless when

freshly prepared, but the solution soon turns brown owing to the oxidation of the hydrogen iodide: $4\text{HI} + \text{O}_2 = 2\text{H}_2\text{O} + 2\text{I}_2$, and the dissolution of the separated iodine. Gaseous hydrogen iodide is also decomposed when mixed with hydrogen and exposed to light. The easy reduction of hydrogen iodide: $2\text{HI} = \text{I}_2 + \text{H}_2$ corresponds with the energetic reducing qualities of this acid. It is largely used as a reducing agent in organic chemistry. The following experiment illustrates this action:

When a deflagrating spoon containing boiling potassium chlorate is plunged in a jar of hydrogen iodide, clouds of iodine are liberated from what appears to be a large red flame of oxygen burning in hydrogen iodide. Nitric acid vapours led from a hot mixture of sulphuric acid and potassium nitrate into a jar of hydrogen iodide, burn with a red flame; so also does a rapid stream of chlorine. A jar of sulphur dioxide held mouth to mouth in a jar of hydrogen iodide, give a deep red flash, and iodine is deposited on the sides of the jar.

The composition of hydrogen iodide can be determined as indicated for hydrogen bromide. The vapour density is 251.8 ($\text{H}_2 = 2$), and the formula HI.

§ 7. Iodides and Bromides.

Hydriodic acid resembles hydrobromic and hydrochloric acids, and forms salts—iodides. The chlorides likewise resemble the iodides and bromides, and they all can be prepared by similar processes. It is, however, worth noting that many of the so-called insoluble iodides precipitated by adding a solution of potassium iodide to a solution of a metallic salt, are soluble in an excess of a solution of potassium iodide. Most metallic iodides when heated in air furnish the metal or a metallic oxide and liberate free iodine. The iodides are usually less volatile than the corresponding chlorides and bromides. Many iodides have characteristic colours.

Potassium iodide, KI.—This salt, as well as potassium bromide, KBr, and potassium chloride, can be prepared by similar methods, namely, by the action of the corresponding acids upon the hydroxides or carbonates; by the action of the elements on solutions of the hydroxide followed by evaporation to dryness and subsequent ignition to decompose the oxy-salts. The iodide is prepared by the following process (using bromine in place of iodine if the bromide is wanted):

Add 25 grams of iodine in small quantities at a time to a mixture of 50 c.c. of water with 7 grams of iron-turnings in a flask with constant agitation. Warm the mixture until the iodine has formed a yellow solution of ferrous iodide, FeI_2 . Decant off the clear solution and mix it with 5 grams more of iodine in order to convert the ferrous salt to ferric iodide. Warm the mixture until all the iodine is dissolved and pour it into a boiling solution of 17 grams of potassium carbonate in 50 c.c. of water. The precipitate becomes flocculent after it has been heated for some time. Test the clear solution to make sure that it is free from iron, if not, add more potassium carbonate to the boiling solution. Evaporate the clear solution for cubic crystals of potassium iodide.

Potassium iodide and bromide are readily soluble in water, and, like the chloride, crystallize in cubes: 100 c.c. of water dissolve 35 grams of potassium chloride; 65, potassium bromide; and 144, potassium iodide at 20°. Potassium iodide and bromide are used in medicine and in photography.

Silver iodide, AgI.—This salt is formed by dissolving silver in concentrated hydriodic acid, or by treating silver nitrate with a soluble iodide. Silver iodide is a yellow crystalline solid; it absorbs gaseous ammonia,

forming a compound 2AgI.NH_3 , which decomposes on exposure to air into ammonia and silver iodide. Silver bromide, AgBr , is a pale yellow solid formed like the iodide with hydrobromic acid, etc. It does not absorb gaseous ammonia like the chloride and iodide. Silver bromide is decomposed by chlorine, and at 100° by hydrochloric acid. At ordinary temperatures, hydrobromic acid converts silver chloride into silver bromide. The reaction is reversible: $\text{AgCl} + \text{HBr} \rightleftharpoons \text{AgBr} + \text{HCl}$. When either silver bromide or silver chloride is treated with hydriodic acid, or potassium iodide, silver iodide is formed. This is due to the fact that silver iodide is far less soluble than the other salts, and consequently separates from the sphere of the reaction. The solubilities of these silver salts in water and in ammonia are as follows:

TABLE XIII.—SOLUBILITIES OF SILVER SALTS.

One litre.	AgF grms.	AgCl grms.	AgBr grms.	AgI grms.
Water at 20°	1818	0.0016	0.000084	0.0000028
Ammonia, 5 per cent. .	—	2.3	0.114	—
Ammonia, 10 per cent. .	—	78.4	3.67	0.035

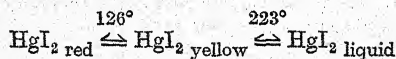
Silver bromide, like silver chloride, is very sensitive to light, and is largely used in dry-plate photography.

Cuprous iodide, CuI .—When aqueous solutions of potassium iodide and a copper salt are mixed together, it is possible that cupric iodide, CuI_2 , is formed: $\text{CuSO}_4 + 2\text{KI} = \text{CuI}_2 + \text{K}_2\text{SO}_4$; but the product is so unstable that it instantly decomposes into almost colourless cuprous iodide and iodine: $\text{CuI}_2 = \text{CuI} + \text{I}_2$ followed by $2\text{I} = \text{I}_2$. The beginning and end of the reaction are thus represented: $2\text{CuSO}_4 + 4\text{KI} = 2\text{CuI} + \text{I}_2 + 2\text{K}_2\text{SO}_4$. This reaction is used in the determination of copper; and for separating iodine from chlorides and bromides, since the two latter salts do not give a cuprous salt under these conditions.

The case of aurous iodide, AuI , is interesting. Meyer made it by shaking precipitated gold with a solution of iodine in potassium iodide for about 24 hrs. Lemon yellow crystals of aurous iodide are formed which decompose quickly on exposure to air: $2\text{AuI} \rightleftharpoons 2\text{Au} + \text{I}_2$. The system is univariant. There are two solid phases—gold and aurous iodide—and one variable—the pressure of iodine vapour. The vapour pressure of iodine must be greater than the dissociation pressure of aurous iodide at, say, 25° , or the latter would dissociate under the conditions of Meyer's experiment. Hence, when metallic gold is brought in contact with iodine, the two unite until the pressure of iodine has reached its equilibrium value 0.943 of that of pure iodine; and conversely, when aurous iodide is confined in space, it will dissociate until the equilibrium pressure of the iodine has been attained. Similar results were obtained with solutions of iodine in carbon tetrachloride in which gold and aurous iodide are not soluble. Aurous iodide is precipitated when potassium iodide is added to an excess of auric chloride. Aurous bromide, AuBr , is made by heating auric bromide. Its properties are similar to those of the chloride. It decomposes in a dry atmosphere at

ordinary temperatures. Auric bromide, AuBr_3 , decomposes into aurous bromide at ordinary temperatures. Hence, the end-reaction $2\text{AuBr} = 2\text{Au} + \text{Br}_2$ proceeds by the intermediate formation of auric bromide which in turn decomposes into aurous bromide and bromine.

Mercuric iodide, HgI_2 .—When a solution of mercuric chloride, HgCl_2 , is treated with a solution of potassium iodide, a yellow precipitate is obtained which changes in a few seconds to a scarlet colour. The scarlet iodide is also made by rubbing 2 parts of mercury with 2.54 parts of iodine wetted with alcohol in a mortar. Mercuric iodide is but sparingly soluble in water, but it is fairly soluble in alcohol and in nitric acid. It readily dissolves in an excess of mercuric salt, and in an excess of potassium iodide. A solution of mercuric iodide in potassium iodide, made alkaline with potassium or sodium hydroxide, is called Nessler's solution. Nessler's reagent is considered to be a solution of the complex potassium mercuric iodide, K_2HgI_4 . It gives a yellow or brown coloration in the presence of ammonia, due to the formation of oxydimercuric ammonium iodide, $\text{NH}_2\text{Hg}_2\text{OI}$. The intensity of the colour depends on the amount of ammonia present. This coloration is a delicate test for ammonia. Mercuric iodide is dimorphous. If the red tetragonal crystals be heated above 126° , they change into yellow rhombic crystals, which reform the original red iodide on cooling. If the yellow iodide persists at ordinary temperatures, it rapidly changes to the scarlet iodide when rubbed with a glass rod.



Mercuric iodide melts at 223° to a red liquid, and a part sublimes, forming yellow rhombic needles. Mercurous iodide, HgI , is formed when mercuric iodide or iodine is rubbed up with the right proportion of mercury in the presence of a little alcohol. It is also formed when a soluble mercurous salt, say, mercurous nitrate, is mixed with a solution of potassium iodide. It is a greenish yellow coloured powder slightly soluble in water. The solution decomposes on standing, particularly if heated, forming mercury and mercuric iodide. The two iodides crystallize in the same system and the crystal angles are nearly the same; hence, they satisfy one test of isomorphism, but not the other tests.

§ 8. Calcium Fluoride.

Calcium fluoride, fluorspar, or fluorite occurs in veins very frequently associated with lead ores, and sometimes by itself. In the limestone caves of the Matlock district, Derbyshire, veins of fine crystals of this mineral are exposed. In Derbyshire the mineral is called "blue john" or "Derbyshire spar." Fluorspar also occurs in many other districts. The crystals belong to the cubic system, and the mineral occurs in cubes, octahedra, and related forms. The crystals may be colourless, or tinted red, brown, yellow, green, blue, or violet by traces of contaminating metallic oxides, etc. Some of the crystals are very pleasing, and such are used for ornamental purposes, jewellery, etc. The crystals are fluorescent. The fluorescence takes the form of a deep blue light when, say, purple fluorspar is exposed to light. A good illustration of fluorescence

is furnished by a dilute solution of fluorescein in very dilute ammonia. This solution is pale orange in transmitted light, but by reflected light it glows with a faint green fluorescence; a colourless solution of quinine sulphate has a blue fluorescence, and uranium glass a green fluorescence. It is supposed that the fluorescence is produced by the absorption of light radiations which set the atoms or molecules vibrating with a period of vibration different from that of the incident rays, and the rays emitted by the vibrating particles accordingly have a different period from that of the incident rays. In other words, the body absorbs rays of one wave-length and transmits them into rays of a different wave-length. This is illustrated by holding a beaker of a solution of quinine sulphate in the ultra-violet invisible spectrum when the solution glows with a blue fluorescence. A more vivid effect is obtained by exposing a piece of cardboard smeared with barium cyanoplatinate to ultra-violet invisible light. No effect is observed in ultra-red light.

Small fragments of fluorspar crystals become luminous or fluorescent with a greenish or orange light when warmed on a hot plate—*thermoluminescence*; or if the crystals be ground in a mortar and pestle—*Triboluminescence*—eddies of light follow the track of the pestle. It has been proposed to mix this mineral with the carbon used for arc lighting so as to increase the luminosity of the arc light and decrease current consumption. When heated to about 902° , fluorspar melts to an opaque greyish-white enamel. This property of fluorspar was mentioned by G. A. Agricola in 1529, and he called the mineral "fluor lapis," literally "fluxing stone"—from the Latin *fluere*, to flow. The German miners apply the term "spath" to all transparent or translucent minerals with a well-marked cleavage, hence the German term for this mineral—"Flussspath." Fluorspar is used as a flux in metallurgy; and in the manufacture of glass, enamels, and glazes.

No gas is evolved when fluorspar is melted, but if fluorspar be strongly heated in an oxidizing flame, on charcoal, a pungent acrid smell, resembling hydrogen chloride, can be detected by bringing the nose near to the charcoal. The fumes redden blue litmus, and if the residue be moistened with water and tested with red litmus, the paper will turn blue. When fluorspar is mixed with sulphuric acid in a test-tube, no perceptible action occurs, but if the mixture be heated, a gas is given off and the glass is strongly corroded, showing that the gas developed by the action of hot sulphuric acid on fluorspar decomposes glass. The gas also attacks porcelain, zinc, copper, silver, etc., but it does not act very markedly upon lead, gold, platinum, wax, paraffin, and rubber. The three latter will not stand heating, and consequently the further investigation of this gas—the "spirit of fluorspar"—hydrogen fluoride, must be conducted in lead, gold, or platinum vessels.

§ 9. Hydrogen Fluoride—Hydrofluoric Acid.

Molecular weight (over 90°), HF = 20.01. Melting point, -92.3° ; boiling point, 19.4° . Specific gravity of the liquid at 12.8° , 0.988. Vapour density ($H_2 = 2$) over 90° , 20.58.

Manufacture of hydrofluoric acid.—An aqueous solution of hydrofluoric acid is manufactured in the following manner: The best quality

of powdered fluorspar, free from silica, is gently heated to about 130° with concentrated sulphuric acid—containing 10 per cent. of water—in a cast-iron pot with a cast-iron cover dipping into an annular trough and sealed with concentrated sulphuric acid. The reaction is represented: $\text{CaF}_2 + \text{H}_2\text{SO}_4 = \text{CaSO}_4 + 2\text{HF}$. Some *calcium fluosulphonate*, $\text{Ca}(\text{FSO}_3)_2$, is found in the retort. The cast-iron retorts are provided with a series of leaden boxes as condensers. These contain water or dilute hydrofluoric acid to absorb the gas from the retorts. The condensers are submerged in water to keep them cool, and the acid so obtained is collected in leaden bottles. Up to 80 per cent. of the theoretical yield can be obtained; with a more concentrated acid the yield is less; with sulphuric acid containing about 50 per cent. of sulphur trioxide, *fluosulphonic acid*, HFSO_3 , distils over; and with sulphuric acid containing 60 per cent. of sulphur trioxide, almost pure fluosulphonic acid is formed. Hydrofluoric acid is placed on the market in leaden, guttapercha, or wax (cerasine) bottles.

Hydrogen fluoride.—If potassium carbonate be neutralized with hydrofluoric acid, and then evaporated to dryness, potassium fluoride separates in cubic crystals readily soluble in water. If the potassium fluoride be dissolved in hydrofluoric acid and evaporated, crystals of the double salt potassium hydrogen fluoride— $\text{KF} \cdot \text{HF}$ —are obtained. This double salt is also called “acid potassium fluoride,” “potassium bifluoride,” and also **Fremy's salt**. If this salt be dried by fusion in a platinum retort, and the platinum retort be then connected with a long platinum tube and platinum bottle immersed in a freezing mixture, the double fluoride decomposes when heated to redness: $\text{KHF}_2 = \text{KF} + \text{HF}$, and the anhydrous hydrogen fluoride is condensed to a liquid. Anhydrous hydrogen fluoride can also be prepared by passing dry hydrogen over dry silver fluoride, and cooling the products of the reaction so as to condense the hydrogen fluoride to a liquid.

Properties.—Anhydrous hydrogen fluoride is a limpid liquid which fumes strongly in air. It is very poisonous and dangerous to manipulate. It forms ulcerated sores if a drop comes in contact with the skin. P. Louyet, about 1850, and J. Nicklès, of Nancy, died in 1869 from accidentally breathing the vapour of this acid while trying to isolate fluorine. The metals potassium and sodium dissolve in the pure acid, forming the corresponding fluorides and hydrogen. The liquid acid boils at 19.5° , and freezes at -102.5° . The crystals melt at -92.3° . Hydrogen fluoride is very soluble in water, forming a corrosive liquid which readily dissolves many metals with evolution of hydrogen: $\text{Fe} + 2\text{HF} = \text{FeF}_2 + \text{H}_2$. Silver and copper also dissolve in the acid. If the acid be more concentrated than about 43 per cent. HF , it will become weaker on boiling; and if more dilute, the acid becomes stronger on boiling until an acid containing about 43 per cent. of HF , boiling at 111° (750 mm.), is formed. This distils unchanged in composition.

Etching glass.—Silicon burns in hydrogen fluoride, forming a mixture of gases, silicon fluoride SiF_4 , and hydrogen. Hydrogen fluoride, unless thoroughly dried, attacks quartz and siliceous substances, glass, etc., also forming silicon fluoride: $\text{SiO}_2 + 4\text{HF} = \text{SiF}_4 + 2\text{H}_2\text{O}$. Hydrofluoric acid is used in the analysis of silicates. When most silicates are repeatedly evaporated with hydrofluoric and sulphuric acids, all the silica is volatilized

as silicon fluoride, etc., and the residue of sulphates can be examined by the standard methods. One of the most important properties of hydrofluoric acid is its etching action on glass. Glass may be etched with the gas or with an aqueous solution of the gas. In the former case, the etching appears opaque and dull; in the latter case, shining and transparent. For etching, the glass is covered with a film of wax, and the design to be etched on the glass is drawn on the waxed surface with a stylus; or else the parts of the glass not to be etched are coated with a resistant varnish. The surface is exposed to the action of the acid or gas, and very soon the glass is etched. The wax or varnish is then washed off with turpentine. The corrosive action of the hydrogen fluoride is due to the ready decomposition of the glass in contact with hydrogen fluoride. The silica forms silicon fluoride. The process is used for marking the scales on glass instruments.

Etching tests for fluorides.—In testing for fluorides, the substance under examination is warmed with sulphuric acid in a leaden vessel covered with a watch-glass. The watch-glass is coated with wax, and a design \boxplus is scratched on the wax with a pin or knife, so as to expose the glass to the action of the acid. The wax is afterwards removed, and if the design is etched on the glass, fluorides were present.

Composition of hydrogen fluoride.—G. Gore (1870) measured the volume of hydrogen required to form hydrogen fluoride when heated with silver fluoride. He found that 100 volumes of hydrogen furnished nearly 200 volumes of hydrogen fluoride. The deviations from the ideal volume relations just indicated were attributed by Gore to the "partial condensation of the hydrogen fluoride to the liquid state." Representing the reaction by: $H_2 + 2AgF = 2HF + 2Ag$, he accordingly inferred that hydrogen fluoride contains half its own volume of hydrogen, and half its own volume of fluorine. The formula is therefore H_nF_n . The vapour density at 100° corresponds with the molecule HF; the vapour density at lower temperatures than 80° shows that the molecule polymerizes.

Vapour density.—The vapour density of hydrogen fluoride at 21.4° ($H_2 = 2$) is 51.18 (or 1.773, air = 1); and it diminishes rapidly as the temperature rises, until, at 90° , it is 20.58. This is illustrated by the graph, Fig. 127. The lower number corresponds with a molecular weight HF. Hence, at 90° , hydrogen fluoride contains two atoms. At lower temperatures the molecule polymerizes to H_nF_n . There is, however, nothing to show what the molecules are. They may be partly HF, partly H_2F_2 , H_3F_3 , etc. The facts only permit the statement that below 90° gaseous hydrogen fluoride is a mixture of molecules H_mF_m , H_nF_n , . . . where the values of m , n , . . . are unknown. Similar results are obtained by lowering the pressure, keeping the temperature constant, at, say, 32° . The effect of hydrogen fluoride on the freezing point of water corresponds with the molecule H_2F_2 . If sodium fluoride has the formula Na_2F_2 , the acid is dibasic.

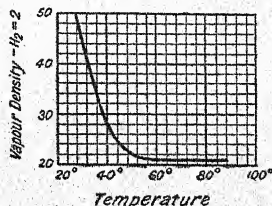
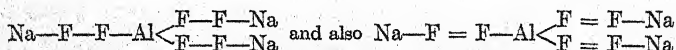
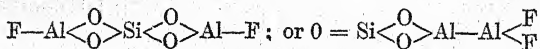


FIG. 127.—Vapour Density of Hydrogen Fluoride.

Fluorides.—The fluorides are made by the action of hydrofluoric acid on the metals, metallic oxides, hydroxides, carbonates, etc. Most of the fluorides are soluble in water. Silver fluoride, AgF , for instance, dissolves in about half its weight of water, and in this respect differs from the corresponding chloride, bromide, and iodide. The solution of silver fluoride in water is alkaline to litmus; and hydrates $\text{AgF} \cdot \text{H}_2\text{O}$, $\text{AgF} \cdot 2\text{H}_2\text{O}$, and $\text{AgF} \cdot 4\text{H}_2\text{O}$ separate as crystals when the solution is concentrated respectively at 30° , 20° , and 10° . Each hydrate dried *in vacuo* over sulphuric acid gives the anhydrous salt. Silver subfluoride, Ag_3F , is said to be formed by the action of a supersaturated aqueous solution of the fluoride on silver. Some consider the alleged silver subfluoride to be a mixture of silver, silver fluoride, and silver oxide. The fluorides of the alkaline earths—calcium, barium, and strontium—are very sparingly soluble, so also is yttrium fluoride. The fluorides also usually unite with hydrogen fluoride, forming the so-called acid fluorides, *e.g.* potassium hydrogen fluoride previously described. The graphic formula may be $\text{K}-\text{F}-\text{F}-\text{H}$, or $\text{K}-\text{F}=\text{F}-\text{H}$. Similarly, the fluorides also unite with one another, forming double salts, *e.g.* the double fluoride of aluminium and sodium, $\text{AlF}_3 \cdot 3\text{NaF}$ or cryolite. The graphic formula for cryolite has been written:



where fluorine is supposed to be either bi- or ter-valent. This curious property of fluorine is also illustrated by the polymerization of hydrogen fluoride at ordinary temperatures. In the topazes, fluorine and hydroxyl appear to be isomorphous for the relative proportion $\text{F}:\text{OH}$ in these minerals varies very much:



the fluorine may be replaced by OH . Several other formulæ have been proposed.

Aluminium fluoride, AlF_3 .—The anhydrous fluoride is made by the action of gaseous hydrogen chloride upon a mixture of calcium fluoride and alumina heated white hot in a graphite tube. The aluminium fluoride volatilizes, and calcium chloride remains behind: $3\text{CaF}_2 + \text{Al}_2\text{O}_3 + 6\text{HCl} = 3\text{H}_2\text{O} + 3\text{CaCl}_2 + 2\text{AlF}_3$. It is also made by the action of silicon fluoride, SiF_4 , upon alumina. The crystalline hydrate, $2\text{AlF}_3 \cdot 7\text{H}_2\text{O}$, is made by dissolving alumina or the metal in aqueous hydrofluoric acid. As indicated above, aluminium fluoride combines with alkaline fluorides, forming double salts. Thus, sodium aluminium fluoride, $\text{AlF}_3 \cdot 3\text{NaF}$, is made by digesting aluminium hydroxide, $\text{Al}(\text{OH})_3$, with sodium fluoride. The salt occurs native in South Greenland as a white, glassy, crystalline solid which resembles clouded ice in appearance, hence the name *cryolite*, literally "ice stone"—from the Greek *κρύος* (kryos), ice; *λίθος* (lithos), stone. Cryolite melts at about 977° , and it is used as a flux. It is also used in the manufacture of alum and aluminium salts, sodium salts, hydrofluoric acid, and the fluorides. Cryolite is not now used as a source of aluminium metal.

§ 10. Fluorine.

Atomic weight, $F = 19$; molecular weight, $F_2 = 38$; uni- or tri-valent. Melting point, -223° ; boiling point, -187° ; vapour density ($H_2 = 2$), 37.7; air = 1) 1.31; specific gravity of liquid, 0.988 at 12.8° .

Occurrence.—Fluorine does not occur free in nature. There is, however, a case recorded by H. Moissan where free fluorine occurs as an inclusion in crystals of fluorspar from Quincié. The compounds of fluorine are widely distributed, though not abundantly, in such minerals as cryolite, fluorspar, etc. Small quantities occur in some of the micas, topaz, tourmaline, etc. It is also said to occur in all rocks, thermal waters, and vapours coming from beneath the earth's crust. Traces occur in seawater, some mineral springs, bones, teeth, blood, milk, plants, etc. The brain of man has about 3 mgrms., and some physiologists believe that the presence of fluorine is necessary in some subtle way, in order that the animal organism can assimilate phosphorus.

History.—The fact that glass is attacked when exposed to the fumes produced when fluorspar is warmed with sulphuric acid was known to Schwankhard in 1670, and in 1771 K. W. Scheele stated that fluorspar is the calcium salt of "a peculiar acid"—*fluoric acid*. He prepared this acid by heating fluorspar with sulphuric acid in a tin retort. J. L. Gay-Lussac and J. Thénard (1807) prepared anhydrous hydrogen fluoride, and, following Lavoisier's school, considered fluoric acid to be a compound of water with the oxide of a new element—"fluorium." In 1810 A. Ampère wrote to H. Davy suggesting "many ingenious and original arguments" in favour of the analogy between hydrochloric and hydrofluoric acids. Ampère concluded that hydrofluoric acid contained no oxygen. The close analogies between calcium iodide, bromide, chloride, and fluoride suggested to A. Ampère that fluorspar is a compound of calcium with an unknown element belonging to the same family as chlorine, bromine, and iodine. Ampère's ideas about hydrofluoric acid were established by H. Davy's experimental work, 1813, and the unknown element was named "fluorine" by analogy with chlorine. No one doubted the existence of fluorine, although it successfully withstood every attempted method of isolation, and for over seventy years had never been seen or handled. Fluorine appeared to be so very powerful that no vessel seemed to be capable of resisting its chemical action; and it was compared with the alkaliest or the universal solvent of alchemy. The brothers Knox sagaciously tried to get fluorine by treating silver or mercury fluoride, in a vessel made of fluorspar itself. Among the many unsuccessful attempts that have been made to isolate this element there are the electrolysis of hydrofluoric acid (H. Davy); electrolysis of fused potassium fluoride (E. Fremy, 1856); the action of chlorine on silver fluoride (H. Davy), and on mercuric fluoride in fluorspar vessels (G. J. and T. Knox, 1836); heating iodine with silver fluoride (H. Kämmerer, 1862); heating silver fluoride (H. Davy); the electrolysis of liquid fused silver fluoride (G. Gore, 1869); heating the unstable uranium fluoride, UF_5 , in oxygen (H. B. Dixon); the action of oxygen on fused calcium fluoride (E. Fremy, 1856); heating lead fluoride, PbF_4 , and also cerium fluoride, CeF_4 (B. Brauner, 1881); etc. About 1884 H. Moissan unsuccessfully tried to isolate it by sparking the gaseous fluorides of arsenic, phosphorus, boron, silicon; by the action

of incandescent platinum on the fluorides of phosphorus and silicon; and the electrolysis of arsenic trifluoride. The feat was accomplished in 1886, when H. Moissan isolated the gas by the electrolysis of a solution of potassium fluoride in liquid hydrogen fluoride, and thus solved what H. E. Roscoe called, "one of the most difficult problems in modern chemistry."

Preparation.—When an electric current is passed through a concentrated solution of hydrogen chloride, chlorine is liberated at the anode, and hydrogen at the cathode. When aqueous hydrofluoric acid is treated in the same way, water alone is decomposed, for oxygen is liberated at the anode, and hydrogen at the cathode. The anhydrous acid does not conduct electricity, and it cannot therefore be electrolyzed. Moissan found that if potassium fluoride be dissolved in the liquid hydrogen fluoride, the solution conducts electricity, and when electrolyzed, hydrogen is evolved at the cathode, and fluorine at the anode. The primary products of the electrolysis are fluorine at the anode, potassium at the cathode: $2\text{KHF}_2 = 2\text{HF} + 2\text{K} + \text{F}_2$. The potassium reacts with the hydrogen fluoride

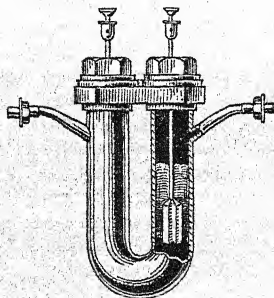


FIG. 128.—Tube for the Electrolysis of Hydrofluoric Acid.

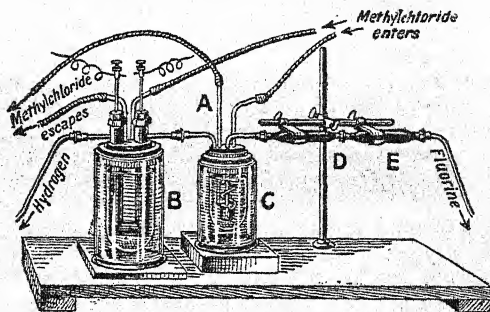


FIG. 129.—Moissan's Process for Fluorine.

reforming potassium fluoride and liberating hydrogen: $2\text{K} + 2\text{HF} = 2\text{KF} + \text{H}_2$.

The electrolysis was first conducted in a U-tube made from an alloy of platinum and iridium which is less attacked by fluorine than platinum alone. Later experiments showed that a tube of copper could be employed. The copper is attacked by the fluorine, forming a surface crust of copper fluoride which protects the tube from further action. Electrodes of the platinum iridium alloy are used. A tube is illustrated in Fig. 128. The open ends of the tube are closed with fluorspar stoppers ground to fit the tubes and bored with holes which grip the electrodes. The joints are made air-tight with lead washers and shellac. The U-tube, during the electrolysis, is surrounded with a glass cylinder B, into which liquid methyl chloride is passed from a steel cylinder *via* the tube A, Fig. 129. Liquid methyl chloride boils at -23° , and it escapes through an exit tube. The fluorine is passed through a spiral platinum tube also placed in a bath of evaporating liquid methyl chloride, C. This cools the spiral tube down to about -50° , and condenses gaseous hydrogen fluoride, which escapes with the fluorine from the U-tube. The fluorine then travels through two

platinum tubes, *D* and *E*, containing lumps of sodium fluoride, which remove the least traces of hydrogen fluoride. A glass cylinder is placed outside each of the two cylinders containing methyl chloride. The outer cylinders contain a few lumps of calcium chloride, so as to dry the air in the vicinity of the cold jacket, and prevent the deposition of frost on the cylinders.

Properties.—Fluorine is a light canary-yellow gas which condenses to a clear yellow liquid boiling at -187° ; and freezes to a pale yellow solid melting at -233° ; at -252° the solid is colourless. Fluorine is probably the most active element known. It combines with hydrogen with explosion, even in the dark, and at low temperatures, say, -210° . It decomposes water, forming hydrogen fluoride, and liberates oxygen highly charged with ozone. Sulphur, selenium, and tellurium melt and take fire in the gas, forming a mixture of fluorides. Iodine, bromine, phosphorus, arsenic, and antimony combine with the gas with incandescence; so do crystalline silicon, amorphous boron, powdered charcoal. All metals are acted upon by the gas; some take fire spontaneously—*e.g.* the metals of the alkalis and alkaline earths; others—*e.g.* magnesium, aluminium, nickel, and silver—require warming to start the reaction; others again require heating to, say, 300° , *e.g.* gold and platinum. Fluorine also liberates chlorine from sodium chloride and from carbon tetrachloride, CCl_4 . Liquid fluorine has no action on silicon, phosphorus, sulphur, and glass. Fluorine is one of the few elements which is not known to form an oxide. It does not react with nitrogen, oxygen, or chlorine, even at the temperature of the electric arc or induction discharge. Palladium and iridium are attacked at about 500° , and rhodium scarcely at all. Glass is attacked by hydrofluoric acid at ordinary temperatures; but fluorine attacks glass so slowly that the gas can be preserved in sealed glass tubes for some time.

No compound of fluorine with chlorine is known. Bromine forms bromine trifluoride, BrF_3 ; and iodine forms iodine pentafluoride, IF_5 . Chlorine can unite with iodine, but there is no reliable evidence of the existence of compounds of bromine with chlorine. The known compounds of iodine with chlorine are: iodine monochloride, ICl ; and iodine trichloride, ICl_3 ; and with bromine: iodine monobromide, IBr . Higher iodine bromides have been reported, but their existence has not been clearly established. A similar remark applies to the so-called bromine monochloride.

Atomic and molecular weight of fluorine.—The combining weight of fluorine has been established by converting calcium fluoride, potassium fluoride, sodium fluoride, etc., into the corresponding sulphates. J. B. A. Dumas (1860) found that 1 gram of pure potassium fluoride furnishes 1.4991 grams of potassium sulphate. Given the atomic weights of potassium 39.1, sulphur 32.07, oxygen 16, it follows that if x denotes the combining weight of fluorine with 39.1 grams of potassium, $1 : 1.4991 = 2\text{KF} : \text{K}_2\text{SO}_4 = 2(39.1 + x) : 174.27$; or, $x = 19$. The best determinations range between 18.97 and 19.14, and the best representative value of the combining weight of fluorine is taken to be 19. No known volatile compound of fluorine contains less than 19 parts of fluorine per molecule, and accordingly this same number is taken to represent the atomic weight. The vapour density of fluorine is 1.31 (air = 1), that is,

$28.755 \times 1.31 = 37.7$ ($H_2 = 2$). The molecule of fluorine is therefore represented by F_2 .

Questions.

1. Potassium iodide is liable to contain potassium iodate, and calomel is liable to contain corrosive sublimate. How do you account for the presence of these impurities, and how would you test for their presence?—*St. Andrews Univ.*

2. Name two minerals containing fluorine, and write their formulæ. How and by whom was fluorine first isolated? How is hydrofluoric acid prepared, in what form is it usually used in the laboratory, and how is it employed in etching?—*Princeton Univ., U.S.A.*

3. By what method is hydrogen fluoride prepared in a state of purity? Contrast its properties with those of the hydrides of chlorine, bromine, and iodine. How has fluorine been isolated?—*London Univ.*

4. How would you prepare a specimen of pure hydrogen iodide? Give examples of its reducing action.—*St. Andrews Univ.*

5. What is meant by the term "catalytic agent"? Describe the use of such an agent in the preparation of hydrogen bromide. What is the chief source of iodine at the present time?—*Cornell Univ., U.S.A.*

6. What do you understand by a reversible chemical action? Cite examples. Point out the conditions affecting the course of the action.—*St. Andrews Univ.*

7. Explain the nature, from a chemical point of view, of the chief operations involved in the production of a photograph.—*London Univ.*

8. Describe the preparation and properties of hydrogen bromide and hydrogen iodide. Why are these gases not commonly prepared by reactions similar to that used in the ordinary preparation of hydrogen chloride?—*Victoria Univ., Manchester.*

9. Describe the effects observed when chlorine water is added (a) to mercurous chloride, (b) to potassium iodide solution. What inference do you draw about the reactions that have taken place? Why?—*Sheffield Scientific School, U.S.A.*

10. Define and give examples of thermal dissociation, kinetic equilibrium, reversible reaction, electrolysis and reduction.—*Princeton Univ., U.S.A.*

11. What experiments have been made with the object of isolating fluorine? and how do you account for their failure?—*London Univ.*

12. Describe and explain the appearances observed when (a) sodium chloride, (b) sodium bromide, (c) sodium iodide is heated with concentrated sulphuric acid. What operations would be necessary to cause hydrogen to combine with chlorine, bromine, and iodine respectively?—*London Univ.*

13. The rate of chemical change may be altered by (a) temperature, (b) catalysis, (c) solution. Describe accurately one experiment illustrating the change in rate of a reaction, which may be brought about by each of these factors.—*London Univ.*

14. Describe a convenient laboratory method for preparing chlorine. How would you dry and collect the gas? What action has chlorine on (a) antimony, (b) an aqueous solution of potassium bromide, (c) an aqueous solution of potassium hydroxide?—*Victoria Univ., Manchester.*

15. Sketch the apparatus you would employ and explain, with all essential practical details, the method you would adopt to prepare a saturated solution of hydriodic acid.—*Board of Educ.*

16. Explain the following quotation: "Scheele rightly regarded manganese (i.e. manganese dioxide) as a substance with a greed for phlogiston, and considered that it removes this principle from muriatic acid, which thereby acquires the properties which characterize dephlogisticated muriatic acid, and that this gas reacts on bodies containing phlogiston by divesting them of this principle" (Berthollet, 1785). In what way did Lavoisier's theory of acid composition support the above reasoning?—*Madras Univ.*

17. J. J. Berzelius found that 100 parts by weight of fluorspar when heated with an excess of sulphuric acid gave 175 parts of anhydrous calcium sulphate. Elemental fluorine was not known at that time, and yet the combining weight of fluorine was calculated from this result to be nearly 19. How was this done? What considerations would lead you to assign the formula HF and CaF_2 to the fluorides of hydrogen and calcium respectively?—*Science and Art. Dept.*

18. How would hydrogen iodide heated in a closed vessel be represented by the phase rule?

CHAPTER XV

THE OXIDES AND OXYACIDS OF CHLORINE, BROMINE, AND IODINE

§ 1. Chlorine Monoxide.

Molecular weight, $\text{Cl}_2\text{O} = 86.92$. Boiling point, $+3.8^\circ$. Relative vapour density ($\text{H}_2 = 2$), 86.5 ; (air = 1) 3.01 .

Preparation.—This compound is prepared by the action of chlorine gas on mercuric oxide. A slow current of dry chlorine from the apparatus *AB*, Fig. 130, is passed through a glass tube *C* containing dry precipitated mercuric oxide previously heated for about an hour to about 400° . The tube is cooled by immersion in cold water, and the issuing gas is passed through a U-tube *D* cooled with ice and salt. Brownish-yellow mercuric oxychloride and chlorine monoxide are formed: $2\text{HgO} + 2\text{Cl}_2 = \text{Hg}_2\text{OCl}_2 + \text{Cl}_2\text{O}$. The gas condenses in the U-tube. If freshly precipitated mercuric oxide be used, the chlorine acts too vigorously, forming mercuric chloride and liberating oxygen—maybe explosively; if the mercuric oxide be in coarse fragments, the reaction is too slow.

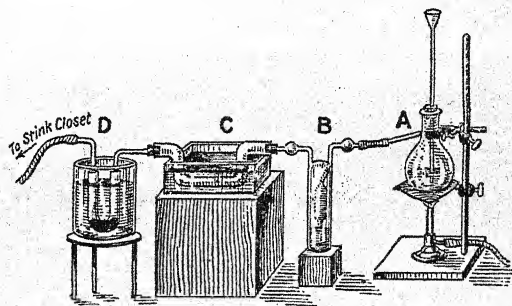


FIG. 130.—Preparation of Chlorine Monoxide.

Properties.—

Chlorine monoxide is a pale orange-yellow gas with the greenish tinge characteristic of chlorine. Its smell is not unlike (but is easily distinguishable from) that of chlorine. One volume of water dissolves 200 volumes of the gas at 0° ; and the solution behaves like an acid—hypochlorous acid, $\text{HClO} : \text{Cl}_2\text{O} + \text{H}_2\text{O} = 2\text{HOCl}$. Hypochlorous acid forms salts—hypochlorites—with the bases. The hypochlorite radicle " ClO " then acts as a monad. Hence chlorine monoxide is also hypochlorous anhydride. If the gas be passed through a tube surrounded with a mixture of ice and salt, the gas condenses to a reddish-brown liquid between -19° and -20° . The liquid boils at 5° and 738 mm. pressure. Both the liquid and the gas are very unstable. Mere contact of the gas with sulphur, phosphorus, and many carbon compounds—*e.g.* caoutchouc,

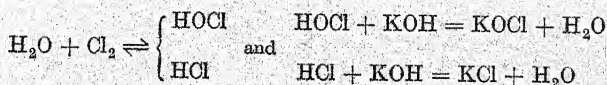
turpentine, etc.—lead to decomposition with explosive violence. Dry calcium chloride is not attacked by chlorine monoxide, but if moisture be present it forms *calcium hypochlorite*, $\text{Ca}(\text{OCl})_2$, and chlorine: $\text{CaCl}_2 + 2\text{Cl}_2\text{O} = \text{Ca}(\text{OCl})_2 + 2\text{Cl}_2$; and the prolonged action of the gas gives *calcium chlorate*, $\text{Ca}(\text{ClO}_3)_2$. The liquid is more liable to explosion than the gas, for it may explode with a slight mechanical shock, *e.g.* when poured from one vessel to another. As indicated above, the liquid can be distilled under reduced pressure; and it may be exposed to sunlight without decomposition in perfectly clean vessels.

Composition.—Pass a stream of the gas through a capillary tube with three bulbs, and gently heat the tube in front of the first bulb. The gas decomposes before it enters the bulbs, the capillary tube prevents an explosion. When the bulbs are filled with the products of decomposition, each bulb can be sealed off and the contents examined. The free chlorine in each bulb is absorbed by potassium hydroxide, the results show that two volumes of chlorine accompany every one volume of oxygen. Since according to Avogadro's hypothesis, equal volumes of these gases contain the same number of molecules, and since both chlorine and oxygen have two-atom molecules, it follows that the molecule of chlorine monoxide has two atoms of chlorine per one atom of oxygen. The vapour density of chlorine monoxide is 86.92 ($\text{H}_2 = 2$). This corresponds with a molecule containing two atoms of chlorine and one atom of oxygen; hence the formula is written Cl_2O . The action of heat is to resolve two volumes of the gas into two volumes of chlorine and one volume of oxygen $2\text{Cl}_2\text{O} \rightarrow 2\text{Cl}_2 + \text{O}_2$.

§ 2. Hypochlorous Acid.

The action of chlorine on cold water.—It is probable that a *cold* aqueous solution of chlorine decomposes, forming a mixture of hydrochloric and hypochlorous acids. For equilibrium: $\text{Cl}_2 + \text{H}_2\text{O} \rightleftharpoons \text{HCl} + \text{HOCl}$. This is evidenced by the fact that chlorine water reacts acid with litmus before it bleaches; and some hypochlorous acid can be separated by distillation. However, the amount of the two acids present when the system is in equilibrium is very small. If one of the products, say hydrochloric acid, be removed, the equilibrium is disturbed and the reaction proceeds in the direction needed to re-establish equilibrium. If freshly precipitated mercuric oxide, for example, be present, the hydrochloric acid reacts with the mercuric oxide, forming mercuric chloride, HgCl_2 . The hypochlorous acid is such a weak acid that it has practically no action on the mercuric oxide. The action of chlorine on water containing mercuric oxide is therefore represented: $\text{HgO} + \text{H}_2\text{O} + 2\text{Cl}_2 \rightleftharpoons \text{HgCl}_2 + 2\text{HOCl}$. Similarly, if calcium carbonate be suspended in the water: $\text{CaCO}_3 + \text{H}_2\text{O} + 2\text{Cl}_2 \rightleftharpoons \text{CaCl}_2 + \text{CO}_2 + 2\text{HOCl}$. If the resulting liquid be distilled, a dilute solution of hypochlorous acid passes over.

The action of chlorine on cold solutions of alkali hydroxides.—A similar action occurs if cold water containing a little potassium hydroxide be treated with chlorine, but both acids are neutralized:



equilibrium is disturbed, and the reaction from left to right is almost complete. The net result of the reaction is represented: $\text{Cl}_2 + 2\text{KOH} \rightleftharpoons \text{KCl} + \text{KOCI} + \text{H}_2\text{O}$. The liquid obtained by saturating a solution of potassium hydroxide with chlorine is called *eau de Javelle*—so named because it was prepared by C. L. Berthollet's process at the chemical works on the Quai de Javelle, a suburb of Paris, in 1792. In 1834 A. J. Balard proved that *eau de Javelle* is a mixture of potassium chloride and hypochlorite. The solution is sometimes used for bleaching purposes. If sodium hydroxide be employed, as suggested by A. G. Labarraque, in 1820, the so-called "chlorinated soda," or *Labarraque's solution*, is obtained. The electrolysis of cold solutions of sodium or potassium chloride furnishes sodium or potassium hydroxide and chlorine. If the products of electrolysis are allowed to intermix, sodium or potassium hypochlorites are formed in a similar manner.

Preparation.—As indicated above, a little hypochlorous acid accompanied by hydrochloric acid is formed when chlorine is dissolved in water. Hypochlorous acid can be obtained by distilling a solution of bleaching powder or *eau de Javelle* with dilute nitric acid. Almost all the hypochlorous acid is liberated by the dilute nitric acid: $\text{KOCl} + \text{HNO}_3 \rightleftharpoons \text{KNO}_3 + \text{HOCl}$, while the chloride is scarcely affected since so little nitric acid is present: $\text{KCl} + \text{HNO}_3 \rightleftharpoons \text{KNO}_3 + \text{HCl}$. To avoid this latter reaction, a weak acid like boric acid is more efficient than nitric acid, because a very great excess of boric acid must be present before appreciable quantities of hydrochloric acid can be set free.

Properties.—Pure hypochlorous acid free from water has not been obtained because the acid is so very unstable. The aqueous solution can be concentrated to a golden-yellow liquid, but only the yellowish solutions containing about 5 per cent. of HOCl can be distilled without decomposition. More concentrated solutions are decomposed on warming into chloric (HClO_3) and hydrochloric acids: $3\text{HOCl} = \text{HClO}_3 + 2\text{HCl}$; followed by the reaction $\text{HCl} + \text{HOCl} = \text{H}_2\text{O} + \text{Cl}_2$, for hypochlorous acid is decomposed by acids, and with hydrochloric acid the chlorine of both acids is set free. Hypochlorous acid is a monobasic acid and forms salts with bases. The radicle "ClO" is a monad, and its compounds with the bases are called hypochlorites. Highly concentrated solutions of sodium hydroxide, saturated with chlorine and evaporated at a low temperature, furnish needle-like crystals of sodium hypochlorite— $\text{NaOCl} \cdot 6\text{H}_2\text{O}$ —contaminated with about 3 per cent. of sodium chloride. Similarly, by the action of chlorine on milk of lime, and further concentrating the resulting solution by the alternate addition of more lime and chlorine, crystals of calcium hypochlorite— $\text{Ca}(\text{OCl})_2$ —have been obtained. The crystals are not deliquescent, and keep well. Hypochlorous acid is so feeble in strength that the carbon dioxide of the air is sufficient to displace the acid from hypochlorites. Note the relation between water, hypochlorous acid, and chlorine monoxide; respectively:



The oxidizing action of hypochlorous acid.—When warmed, hypochlorous acid not only furnishes chloric acid— HClO_3 —as indicated above, but it is also decomposed with the evolution of oxygen: $2\text{HOCl} = 2\text{HCl}$

+ O₂. This reaction is particularly active in sunlight and in the presence of oxidizing agents. Thus with silver oxide: $\text{Ag}_2\text{O} + 2\text{HOCl} = 2\text{AgCl} + \text{H}_2\text{O} + \text{O}_2$. If a little nickel or cobalt nitrate solution be added to water containing hypochlorous acid and the mixture warmed in a flask, oxygen is evolved: $2\text{HOCl} = 2\text{HCl} + \text{O}_2$. The cobalt salt acts as a catalytic agent. The mechanism of the reaction is generally supposed to involve the concurrent reactions corresponding with the transformations of the cobalt oxide: $\text{CoO} \rightarrow \text{Co}_3\text{O}_3 \rightarrow \text{CoO} \rightarrow \text{Co}_2\text{O}_3 \rightarrow \text{etc.}$ When oxygen is prepared by this process, bleaching powder suspended in water is the usual source of the hypochlorous acid. In illustration of the oxidizing action of hypochlorous acid, calcium hypochlorite or bleaching powder may be boiled for some time with a solution of lead acetate, puce coloured lead dioxide, PbO_2 , will be precipitated; and if boiled with a solution of a manganous salt, manganese dioxide, MnO_2 , will be precipitated. More prolonged boiling may give a green solution of calcium manganate, or a pink coloured solution of calcium permanganate, $\text{Ca}(\text{MnO}_4)_2$.

The rapid decomposition of hypochlorous acid in sunlight renders it probable that the action of light on chlorine water results in the formation of the hypochlorous acid by hydrolysis: $\text{Cl}_2 + \text{H}_2\text{O} = \text{HCl} + \text{HOCl}$, is at once decomposed: $2\text{HOCl} = 2\text{HCl} + \text{O}_2$, so that the hydrolysis goes to completion, and leaves, as final products, hydrochloric acid, water, and oxygen.

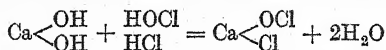
Test.—When an excess of mercury is shaken up with an aqueous solution of hypochlorous acid, a brownish-yellow precipitate of mercuric oxychloride, $\text{HgO} \cdot \text{HgCl}_2$, is formed; this is decomposed by dilute hydrochloric acid—mercuric chloride passes into solution. With chlorine water, mercury gives a white precipitate of mercurous chloride, HgCl ; hence the reaction can be used to distinguish and even to estimate hypochlorous acid in the presence of free chlorine in solution.

§ 3. Bleaching Powder.

History.—The first attempt to apply chlorine to bleaching was made by C. L. Berthollet in 1785. The liquor employed was obtained by the action of chlorine on alkali-lye. Attempts were made to suppress the unpleasant effects of chlorine by mixing it with potash or chalk. Rumours of a mysterious bleaching liquid began to attract attention. In Scotland linen bleaching was practised extensively. The best goods were sent to Holland by the Scotch bleachers to be bleached, and the goods returned to Leith. Chlorine was tried, but disappointment followed. The gas attacked the fabric and the linen rotted. Berthollet's discoveries were communicated to James Watt in Paris about the same time, and Watt soon afterwards brought the news to Scotland. In 1798 Charles Tennant patented a process for the use of cheaper lime in place of potash; at first milk of lime was used, and a year later, solid calcium hydroxide. The patent was later declared void because lime had been used for the same purpose in Lancashire prior to Tennant's patent.

The action of chlorine on calcium hydroxide.—If chlorine be allowed to act upon an aqueous solution of a bivalent base, say calcium hydroxide, in place of potassium or sodium hydroxide, a molecule of each of the two

monobasic acids, formed by the action of chlorine on water, is neutralized by one molecule of the base, and what seems to be a mixed salt is formed :



The salt $\text{Cl}-\text{Ca}-\text{OCl}$ is called "bleaching powder" or "chloride of lime." This substance may also be regarded as a molecular compound of calcium chloride with calcium hypochlorite: $\text{CaCl}_2 \cdot \text{Ca}(\text{OCl})_2$. The constitution of bleaching powder has been much discussed, and the subject is yet far from being definitely settled. It is very probable that bleaching powder contains but little calcium chloride because (1) all the chlorine can be expelled from it by the action of carbon dioxide. This would not be the case if calcium chloride were present, unless some reaction like that assumed by Balard takes place, viz. $\text{CO}_2 + \text{Ca}(\text{OCl})_2 = \text{CaCO}_3 + \text{CaCl}_2 + \text{Cl}_2\text{O}$, followed by $\text{Cl}_2\text{O} + \text{CO}_2 + \text{CaCl}_2 = \text{CaCO}_3 + \text{Cl}_2$; (2) calcium chloride is very deliquescent, bleaching powder is not; (3) calcium chloride is readily dissolved by alcohol, whereas an alcoholic solution of bleaching powder contains but traces of calcium chloride. The two latter objections have been answered by a subsidiary hypothesis that the calcium chloride is united with the hypochlorite to form a double salt, $\text{CaCl}_2 \cdot \text{Ca}(\text{ClO})_2$. (4) However well prepared, bleaching powder always contains an excess of lime; at any rate, a deposit of slaked lime is always found as a residue when bleaching powder is treated with water. Under very favourable circumstances lime can be saturated with no more than $43\frac{1}{2}$ per cent. of available chlorine. The facts correspond with the formula $\text{Ca}(\text{OCl})\text{Cl}$ first suggested, without proof, by W. Odling in 1861. Since the available chlorine in commercial bleaching powder usually ranges between 36 and 38 per cent., it is evident that the calcium hydroxide, $\text{Ca}(\text{OH})_2$, is not completely saturated with chlorine, and that some free calcium hydroxide is present. Hence commercial bleaching powder is best represented as $\text{Ca}(\text{OCl})\text{Cl} + n\text{Ca}(\text{OH})_2$, where n is very nearly one-half, or $2\text{Ca}(\text{OCl})\text{Cl} + \text{Ca}(\text{OH})_2$.

Preparation.—Bleaching powder is made on a large scale by the action of chlorine on slaked lime, $\text{Ca}(\text{OH})_2$. The lime is spread in 3- or 4-inch layers on perforated shelves in a large chamber, and then raked into furrows. Chlorine is led through the chambers. At first the absorption of chlorine is rapid, but it afterwards slows down. The lime is then turned over from time to time so as to expose a fresh surface. After standing for 12 to 24 hours a shower of fine dust lime is blown into the chamber to absorb the excess of chlorine. The amount of chlorine absorbed is never so complete as is represented by the equation: $\text{Ca}(\text{OH})_2 + \text{Cl}_2 = \text{Ca}(\text{OCl})\text{Cl} + \text{H}_2\text{O}$. The commercial value of the bleaching powder depends on the amount of available chlorine it contains. The amount of available chlorine depends upon the method of preparation, temperature, etc. If the temperature be kept between 30° and 40° , a bleaching powder containing about 40 per cent. of available chlorine has been prepared.

Evaluation of bleaching powder.—The process for the determination of the available chloride depends upon the fact that sodium arsenite is oxidized to sodium arsenate by an aqueous solution of bleaching powder. Hence a standard solution of sodium arsenite is added to a known amount

of an aqueous solution bleaching powder until the solution no longer gives a blue coloration with iodized starch paper. This shows that no available chlorine is present. The amount of sodium arsenite used in the experiment is related with the bleaching powder by the following equation: $\text{Ca}(\text{OCl})\text{Cl} + \text{Na}_3\text{AsO}_3 = \text{Na}_3\text{AsO}_4 + \text{CaCl}_2$, and the theoretical amount of $\text{Ca}(\text{OCl})\text{Cl}$ can therefore be readily computed.

The action of water and acids.—If bleaching powder be treated with cold water, it forms a strongly alkaline solution, and insoluble calcium hydroxide remains. It is probable that the action is due to the hydrolysis of the calcium hypochlorite: $2\text{Ca}(\text{OCl})\text{Cl} = \text{Ca}(\text{OCl})_2 + \text{CaCl}_2$; followed by $\text{Ca}(\text{OCl})_2 + 2\text{H}_2\text{O} = \text{Ca}(\text{OH})_2 + 2\text{HOCl}$. If the bleaching powder be treated with a very dilute acid, hypochlorous acid, HOCl , is formed: $\text{Ca}(\text{OCl})\text{Cl} + \text{HCl} = \text{CaCl}_2 + \text{HOCl}$. If an excess of acid be present, the hypochlorous acid is decomposed, forming water and chlorine: $\text{HOCl} + \text{HCl} = \text{H}_2\text{O} + \text{Cl}_2$.

Bleaching powder decomposes when exposed to atmospheric moisture. The carbon dioxide of the air also reacts like a feeble acid as indicated above. Bleaching powder also decomposes slowly when kept in a well-stoppered bottle. When heated with ammonia, nitrogen is obtained: $2\text{NH}_3 + 3\text{Ca}(\text{OCl})\text{Cl} = 3\text{H}_2\text{O} + 3\text{CaCl}_2 + \text{N}_2$; and when boiled with water and a little cobalt salt, oxygen is obtained as indicated above. Thus, hypochlorous acid, chlorine and oxygen can be obtained from bleaching powder.

Bleaching.—In bleaching by eau de Javelle or by bleaching powder, the fabric is steeped in a clear dilute aqueous solution of the bleaching agent—*chemicking*—and then *soured* by steeping in dilute acid. Hypochlorous acid is thus produced, and then free chlorine. The free chlorine does its work within the fibres of the wet fabric as indicated on p. 282. The bleaching action of hypochlorous acid is generally stated to be twice as great as that of the chlorine it contains, supposing the latter were free: $2\text{Cl}_2 + 2\text{H}_2\text{O} = 4\text{HCl} + \text{O}_2$; $4\text{HOCl} = 4\text{HCl} + 2\text{O}_2$. But it must be remembered that two atoms of chlorine are needed to form one molecule of HOCl , since an equivalent amount of HCl is formed at the same time.

Electrolytic bleaching liquid is preferred for some purposes to bleaching powder, although the latter is rather cheaper than the former. The bleaching liquid is made by the electrolysis of solutions of sodium chloride. The chlorine which accumulates about the anode is mixed with a solution of sodium hydroxide which collects about the cathode. The result is an aqueous solution of chlorine containing sodium hydroxide and chloride, i.e. a mixed solution of sodium hypochlorite and chloride. A brisk circulation of the electrolyte is needed to prevent the escape of chlorine.

§ 4. Potassium and Barium Chlorates.

The manufacture of potassium chlorate, KClO_3 .—We have seen that when chlorine is passed into a cold solution of potassium hydroxide, a mixture of potassium chloride and hypochlorite is formed; and when the solution of the hypochlorite is boiled, it decomposes, forming a mixture of potassium chlorate and chloride; otherwise expressed, the hypochlorite solution oxidizes itself: $3\text{KOCl} = 2\text{KCl} + \text{KClO}_3$. The reaction takes

place in aqueous solution with the evolution of 23·8 Cals. Potassium chlorate is likewise obtained when chlorine is passed into a *hot* (70°) aqueous solution of potassium hydroxide: $6\text{KOH} + 3\text{Cl}_2 = \text{KClO}_3 + 5\text{KCl} + 3\text{H}_2\text{O}$; and the two salts—potassium chloride and potassium chlorate—can be separated by fractional crystallization. Potassium chlorate is far less soluble than the corresponding chloride, p. 30. C. L. Berthollet (1786–8) first isolated this salt, although it appears to have been known to J. R. Glauber (1658), who mistook it for saltpetre. The above method of preparation is due to J. L. Gay-Lussac (1818).

Liebig's process of manufacture.—It will be observed that the amount of chlorate obtained from a given amount of potassium hydroxide is small, because one molecule of potassium chlorate is accompanied by five molecules of potassium chloride as by-product. This loss is serious because the potassium hydroxide is relatively expensive. This led J. von Liebig (1842) to modify the process. It is cheaper to substitute a hot solution of slaked lime in water for the potassium hydroxide solution. In that case: $6\text{Ca}(\text{OH})_2 + 6\text{Cl}_2 = \text{Ca}(\text{ClO}_3)_2 + 5\text{CaCl}_2 + 6\text{H}_2\text{O}$. The clear solution of calcium chlorate and chloride is concentrated a little by evaporation, and a slight excess of potassium chloride is added. Potassium chlorate has but one-tenth the solubility of the corresponding calcium salt, and is far less soluble than the two chlorides; hence by a further concentration of the solution, the least soluble potassium chlorate separates: $\text{Ca}(\text{ClO}_3)_2 + 2\text{KCl} = 2\text{KClO}_3 + \text{CaCl}_2$. The potassium chlorate so obtained is purified by recrystallization.

The electrolytic process of manufacture.—The old process of J. von Liebig is now almost displaced by the electrolytic process; indeed, after the introduction of the electrolytic process the price of chlorate soon fell from £60 to £35 per ton, and abundantly justified H. Davy's prediction (1806): "It is not improbable that the electric decomposition of neutral salts in different cases may admit of economical ones." Hot solutions of potassium chloride are electrolyzed. The initial and end stages of the reactions are represented: $\text{KCl} + 3\text{H}_2\text{O} = \text{KClO}_3 + 3\text{H}_2$; but no doubt chlorine is first formed (p. 278); then potassium hypochlorite; and finally potassium chlorate. The sparingly soluble potassium chlorate crystallizes from the solution during the electrolysis, and thus gives trouble. Since 100 c.c. of water at 20° dissolve about 99 grams of sodium chlorate; and 100 c.c. of water, about 7·2 grams of potassium chlorate, it is best to first prepare sodium chlorate by the electrolysis of sodium chloride; and then treat the solution with potassium chloride as in the case of calcium chlorate described above. Potassium chlorate can be readily isolated by fractional crystallization.

The solubility and the action of heat on potassium chlorate have been described previously. Potassium chlorate is used medicinally; in the manufacture of matches, fireworks, and explosives; as an oxidizing agent in chemical processes, for the preparation of small quantities of oxygen; etc.

Barium chlorate, $\text{Ba}(\text{ClO}_3)_2$.—This salt can be prepared by the electrolysis of a solution of barium chloride; or by the action of chlorine on a hot solution of barium hydroxide, or on water containing barium carbonate in suspension. The separation of barium chlorate and barium chloride by fractional crystallization is rather difficult because the two salts are

almost equally soluble in water. It is best to cool the crystallizing solution to 0° , at which temperature 100 parts of a saturated solution contain 16.9 grams of barium chlorate; and 100 grams of a saturated solution contain 24 grams of barium chloride; at 10° , the corresponding solubilities are 21 and 25; and at 20° , 25 and 26.

§ 5. Chloric Acid.

Preparation.—The chlorates, as we have seen, were first made by C. L. Berthollet, and J. L. Gay-Lussac (1814) obtained chloric acid by the following process: A solution of, say, 80 grams of barium chlorate in 170 c.c. of water is treated with an equivalent quantity of sulphuric acid (24.3 grams of H_2SO_4 in 200 c.c. of water), when barium sulphate and chloric acid are formed: $\text{Ba}(\text{ClO}_3)_2 + \text{H}_2\text{SO}_4 = \text{BaSO}_4 + 2\text{HClO}_3$. The clear solution of chloric acid is decanted from the precipitated barium sulphate, and concentrated by evaporation in vacuo over sulphuric acid. Solutions containing more than about 40 per cent. HClO_3 cannot be prepared, because the chloric acid then decomposes spontaneously into free chlorine, oxygen, perchloric acid— HClO_4 —and water. The first action is probably: $2\text{HClO}_3 = \text{HClO}_4 + \text{HClO}_2$; the chlorous acid, HClO_2 , so formed reacts with the chloric acid, forming chlorine peroxide, ClO_2 : $\text{HClO}_2 + \text{HClO}_3 = \text{H}_2\text{O} + 2\text{ClO}_2$; and the chlorine peroxide breaks down into chlorine and oxygen: $2\text{ClO}_2 = \text{Cl}_2 + 2\text{O}_2$, as indicated below. The initial and end products of the completed reaction are represented: $3\text{HClO}_3 = \text{HClO}_4 + \text{H}_2\text{O} + \text{Cl}_2 + 2\text{O}_2$.

Properties.—The concentrated solution of chloric acid so prepared is a colourless viscid liquid with a pungent smell. It readily decomposes on exposure to light. The solution is stable in darkness provided organic matter be absent. Wood, paper, etc., decompose the acid at once—very often with spontaneous combustion. Blue litmus is first reddened and then bleached by the acid. Even in a dilute solution, chloric acid is a powerful bleaching agent. The acid is monobasic, forming a series of salts—chlorates—where ClO_3 acts as a univalent radicle. The anhydride of the acid—chlorine peroxide, Cl_2O_5 —is not known.

The chlorates.—The chlorates are powerful oxidizing agents. An explosion may occur if a chlorate be mixed with organic matter, charcoal, sulphur, etc., and the mixture struck with a hammer, or heated. Hence mixtures of chlorates with such materials must not be ground together with a pestle and mortar. The materials should be ground separately, and then carefully mixed on paper with a feather. Phosphorus in contact with a chlorate may explode spontaneously. Thus if a drop of a solution of phosphorus in carbon disulphide be allowed to fall on a little potassium chlorate, a loud explosion occurs as soon as the carbon disulphide has evaporated. The red phosphorus in the mixture on the side of a box of safety matches gives a series of sparks when a piece of potassium chlorate is drawn across, and serious accidents have resulted by the accidental rubbing of tabloids of potassium chlorate on the side of a matchbox in the same pocket. The oxidizing action of potassium chlorate in neutral or slightly acid solutions is greatly stimulated by traces of osmium tetrachloride, OsCl_4 , as catalytic agent.

The chlorates are all soluble in water. Potassium chlorate is one of

the least soluble of the chlorates. The salts are fairly stable, but decompose into chlorides and oxygen when heated. The chlorates are recognized by giving no precipitate with silver nitrate; although, after ignition, the silver nitrate will give a precipitate of silver chloride with an aqueous solution of the residue. If a few drops of a solution of indigo sulphate be added to an aqueous solution of a chlorate, and the liquid be acidified with sulphuric acid, and sulphurous acid, or a sulphite be added, the chloric acid is reduced to a lower chlorine oxide which bleaches the blue colour of the indigo. If 3 or 4 drops of concentrated sulphuric acid be allowed to run down the side of a test-tube containing a little chlorate solution mixed with aniline sulphate, a blue colour appears where the two liquids meet. Nitrates and iodates give a brownish red coloration under similar circumstances.

Composition.—The composition of chloric acid was established by J. S. Stas' analyses of silver chlorate. A known amount of silver chlorate was reduced to the chloride by the action of sulphurous acid. Previous analyses had established the exact composition of silver chloride. Stas found that 100 grams of silver chlorate furnished :

Silver chloride—AgCl	74·9205 grams.
Oxygen	25·0795 „
Silver chlorate	100·0000 „

The molecular weight of silver chloride is 143·43; and the atomic weight of oxygen is 16. By division, $74·9206 \div 143·34 = 0·52$; and $25·0795 \div 16 = 1·58$. Hence the ratio of silver chloride to oxygen in the silver chlorate is as 1 : 3; or the empirical formula of silver chlorate is AgClO_3 , and of the acid HClO_3 . The molecular weight of the acid has not been determined satisfactorily.

§ 6. Chlorine Peroxide.

Molecular weight, $\text{ClO}_2 = 67·45$. Melting point, -79° ; boiling point, $9·9^\circ$. Vapour density, $67·29 (\text{H}_2 = 2)$; $2·39 (\text{air} = 1)$.

Preparation.—While studying the action of concentrated sulphuric acid upon potassium chlorate, H. Davy (1815) found that a highly explosive gas

was produced. Finely powdered potassium chlorate is gradually added to concentrated sulphuric acid in a small flask or retort, A, Fig. 131. The salt dissolves, producing a reddish-brown liquid, but no gas is evolved if the liquid be kept cold. When the solution is gradually warmed, by placing the retort in a vessel, B, of warm water, taking care not to heat the glass above the level of the liquid in the retort, chlorine peroxide is evolved as a gas. The first action of the sulphuric acid on potassium chlorate is to form chloric acid: $\text{KClO}_3 + \text{H}_2\text{SO}_4 = \text{KHSO}_4 + \text{HClO}_3$; and the chloric acid is then decomposed into perchloric acid, HClO_4 , chlorine peroxide, and water: $3\text{HClO}_3 = \text{HClO}_4 + 2\text{ClO}_2 + \text{H}_2\text{O}$.

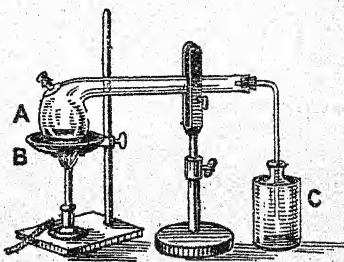
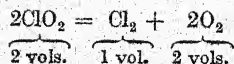


FIG. 131.—Preparation of Chlorine Peroxide.

Properties.—Chlorine peroxide is a reddish-yellow gas with an unpleasant smell which produces headache; very dilute solutions have a smell which is not unpleasant and has been likened to ozone. The gas is much heavier than air, and is collected by the upward displacement of air, *C*, since it decomposes in contact with mercury, and it is fairly soluble in water. Water at 4° dissolves about twenty times its volume of the gas. When the solution is cooled to lower temperatures a crystalline hydrate separates—possibly as $\text{ClO}_2 \cdot 8\text{H}_2\text{O}$. On standing, in the dark, an aqueous solution of chlorine peroxide forms a mixture of chloric and hydrochloric acids. When the gas is cooled, it condenses to a dark red liquid which boils at 9.9°, and the liquid can be frozen to orange-coloured crystals melting at -79°. Chlorine peroxide is very unstable. It decomposes with explosive violence if an electric spark be passed through the gas, or if a hot wire be introduced into the gas. Chlorine peroxide also decomposes into its elements if it be exposed to the light. The gas is liable to suddenly explode, especially if in the liquid or solid condition, or if organic matter be present. Chlorine peroxide is a powerful oxidizing agent—a piece of phosphorus, sugar, or other combustible takes fire spontaneously in the gas.

Two well-known experiments may be cited to illustrate this. Place some crystals of potassium chlorate at the bottom of a test glass, and half fill the vessel with water. A few lumps of yellow phosphorus are dropped into the glass, and concentrated sulphuric acid is allowed to flow through a tube funnel on to the chlorate. The bubbles of chlorine peroxide which are evolved produce bright flashes of light when they come in contact with the phosphorus under the water. Again, powdered sugar and potassium chlorate are mixed with a feather on a sheet of paper and placed on a stone slab. When a drop of sulphuric acid is allowed to fall upon the mass, the chlorine peroxide which is formed ignites the sugar, and the flame rapidly spreads throughout the mass.

Composition.—H. Davy's formula for the gas was confirmed by Pebal in 1875, who obtained a mixture of 1 volume of chlorine with 2 volumes of oxygen, when two volumes of the gas were exploded. In spite of this, the formula was written Cl_2O_4 until Pebal and Schachert, in 1882, found the vapour density of the gas in harmony with ClO_2 . Gay-Lussac determined the composition of chlorine peroxide by passing the gas through a capillary tube with three bulbs of known capacity. The capillary tube was heated before the gas entered the bulbs. Decomposition took place in the capillary tube without explosion. The bulbs therefore contained the decomposition products of the chlorine peroxide—chlorine and oxygen. The chlorine was determined by absorption with potassium hydroxide and the oxygen in a gas-measuring tube. It was found that 100 volumes of the peroxide furnished 100.6 volumes of oxygen and 49.3 volumes of chlorine. Hence it was inferred that



and by the application of Avogadro's hypothesis, that equal volumes contain an equal number of molecules, and that oxygen and chlorine molecules each contain two atoms, it follows that the formula of chlorine peroxide is ClO_2 . This is confirmed by the vapour density—67.29—which corresponds with the molecule ClO_2 , not Cl_2O_4 . The molecular weights of the gas in water, and in carbon tetrachloride, as solvents, agree with

the formula ClO_2 . Thus the chlorine in chlorine peroxide appears, at first sight, to be quadrivalent.

Chlorites and chlorous acid.—When an aqueous solution of chlorine peroxide is treated with an alkali, say, potassium hydroxide, a mixture of potassium chlorite, KClO_2 , and chlorate is formed: $2\text{KOH} + 2\text{ClO}_2 = \text{KClO}_2 + \text{KClO}_3 + \text{H}_2\text{O}$; and if sodium peroxide be employed, sodium chlorite and oxygen are produced—the latter in amounts corresponding with the equation $\text{Na}_2\text{O}_2 + 2\text{ClO}_2 = 2\text{NaClO}_2 + \text{O}_2$. The faintly colourless alkaline solution so obtained is without action on indigo solution, or iodized starch paper, and it becomes yellow when an acid is added. The acidified solution probably contains chlorous acid, HClO_2 . The chlorites are readily decomposed. Lead chlorite $\text{Pb}(\text{ClO}_2)_2$ is precipitated when the solution of sodium chlorite is acidified with acetic acid and treated with lead nitrate. Lead chlorite at 100° decomposes with detonation. The soluble chlorites bleach vegetable colouring matters, even after the addition of arsenious acid. This is not the case with the hypochlorites. Pure chlorous acid has not been prepared; an aqueous solution is obtained by digesting the lead salt with barium carbonate so as to form barium chlorite, $\text{Ba}(\text{ClO}_2)_2$, and finally precipitating the barium as sulphate by $\frac{1}{10}\text{N-H}_2\text{SO}_4$. The solution gradually decomposes: $4\text{HClO}_2 = 2\text{H}_2\text{O} + 3\text{ClO}_2 + \text{Cl}$. The corresponding anhydride—chlorine trioxide, Cl_2O_3 —is unknown.

H. Davy prepared a gas by the action of concentrated hydrochloric acid upon potassium chlorate, which he believed to be chlorine trioxide, with the composition Cl_2O_3 , and which was termed *euchlorine*. *Euchlorine*, however, is a mixture of chlorine with chlorine peroxide in varying proportions.¹ A mixture of potassium chlorate and hydrochloric acid is often employed in the laboratory as an oxidizing agent.

§ 7. Perchloric Acid.

Preparation.—Perchloric acid is formed when chloric acid is heated or exposed to light: $3\text{HClO}_3 = \text{Cl}_2 + \text{HClO}_4 + 2\text{O}_2 + \text{H}_2\text{O}$; and by the action of sulphuric acid on, say, potassium perchlorate: $2\text{KClO}_4 + \text{H}_2\text{SO}_4 \rightleftharpoons \text{K}_2\text{SO}_4 + 2\text{HClO}_4$. Count Stadion first made perchloric acid in 1816. Potassium perchlorate is prepared, as indicated on pp. 204–5, by the action of heat on potassium chlorate.

If 50 grams of pure dry potassium perchlorate be distilled under reduced pressure in a 300 c.c. distilling flask with 150 to 175 grams of concentrated (96 to 97½ per cent.) sulphuric acid, a white crystalline mass of $\text{HClO}_4 \cdot \text{H}_2\text{O}$ collects in the receiver. The water is formed by the decomposition of the acid during the distillation. By the redistillation of this product under reduced pressure, pure perchloric acid can be obtained.

Electrolytic process of manufacture.—Perchloric acid and the perchlorates can be prepared by the electrolysis of the chloric acid and the chlorates respectively. The chlorate is not directly oxidized, as was once

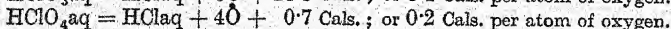
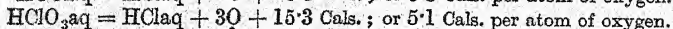
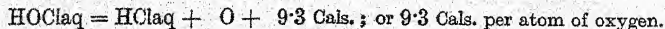
¹ D. I. Mendeléeff thinks that some chlorine trioxide, Cl_2O_3 —vapour density 119—is present because the vapour density of *euchlorine*, according to L. Pebal, is about 80; and the vapour density of chlorine peroxide is 67.3, and of chlorine, 71. Mixtures of these two would give too low a vapour density. Some of the chlorine peroxide, however, may be present as Cl_2O_4 .

supposed. The chlorate ions appear at the anode during the electrolysis: $\text{KClO}_3 = \text{K}^+ + \text{ClO}_3^-$, when the ions are discharged at the electrodes, the potassium reacts with the water forming hydrogen and the chlorate ion forms chloric acid which decomposes as indicated above into perchloric acid and chlorous acids and oxygen: $2\text{ClO}_3 + \text{H}_2\text{O} = \text{HClO}_4 + \text{HClO}_2 + \text{O}$. The oxygen so formed oxidizes the chlorous acid back to chloric acid: $\text{HClO}_2 + \text{O} = \text{HClO}_3$. The initial and end products are therefore represented: $2\text{KClO}_3 + 2\text{H}_2\text{O} = 2\text{KClO}_4 + 2\text{H}_2$. It has been observed empirically that the presence of a trace of potassium dichromate increases the yield.

Properties.—Perchloric acid is a volatile colourless fuming liquid; specific gravity 1.764 at 22°, boils at 14° to 18° under a pressure of 15 to 20 mm., and freezes at -112° to a crystalline solid. If a drop of the acid be brought in contact with paper or wood, instantaneous and violent inflammation occurs; if a drop of the acid be brought in contact with charcoal, a violent explosion occurs. Perchloric acid produces serious wounds in contact with the skin. If the acid be distilled at ordinary pressures, the liquid may gradually become darker and darker in colour, and finally explode violently. The pure acid is also said to be liable to explode after standing some days.

The action of water.—When dropped into water, the two combine with a hissing sound, and the evolution of much heat: $\text{HClO}_4 + \text{Aq} = \text{HClO}_4\text{aq} + 20.3 \text{ Cals.}$ Perchloric acid forms five hydrates containing respectively 1, 2, 2.5, 3, and 3.5 molecules of water. The monohydrate is formed by adding water to the pure acid until crystals begin to appear. This hydrate melts at 50°, and freezes at the same temperature, forming long needle-like crystals. Solutions more concentrated than 71.6 per cent. HClO_4 lose HClO_4 when distilled, and less concentrated solutions lose water. The "constant boiling acid" contains 71.6 per cent. of HClO_4 , boils at 203°, and distils unchanged in composition. It fumes slightly in air, and may be preserved indefinitely, even in light. Perchloric acid slowly volatilizes at 138° without decomposition.

Thermochemistry of the oxychlorine acids.—Perchloric acid is not so powerful an oxidizing agent as chloric acid, and this in turn is less vigorous than hypochlorous acid. This agrees with the stability of the respective acids, and is in accord with the fact that a greater amount of available energy per atom of available oxygen associated with hypochlorous acid than with either chloric or perchloric acid. For instance, the thermochemical equations are represented:



Since the bleaching effect of these reagents is supposed to depend upon the action of nascent oxygen, it follows that in this case there is no particular need for the assumption that atomic oxygen is more active than molecular oxygen, because the "nascent oxygen" is associated with a larger amount of available energy, which can do chemical work. The available energy of each acid must be added to that which free oxygen could give if it alone were performing the same oxidation; according to the principle of maximum work, the amount of energy degraded during a chemical reaction measures the "tendency of a reaction to take place."

Hence the greater oxidizing properties of these acids must, at least in part, be due to the greater amount of available energy associated with their "nascent oxygen" during decomposition.

Perchlorates.—An aqueous solution of perchloric acid reddens litmus, and forms salts—perchlorates—where the radicle " ClO_4 " is univalent. Hence perchloric acid is monobasic. While a dilute aqueous solution of perchloric acid reddens litmus, dilute aqueous solutions of chloric and hypochlorous acids bleach litmus. The potassium salt is one of the least soluble perchlorates. It is practically insoluble in absolute alcohol. When perchloric acid is added to an alcoholic solution of a soluble potassium salt, potassium perchlorate is quantitatively precipitated. The weight of potassium perchlorate so obtained enables the amount of potassium in the given solution to be computed. Unlike the chlorates, the perchlorates are not decomposed by hydrochloric acid; nor do they yield an explosive gas when warmed with concentrated sulphuric acid. They are not reduced to chlorides by sulphur dioxide; and they require a higher temperature for their decomposition than the corresponding chlorates. Potassium perchlorate in aqueous solution is quantitatively reduced to the chloride when boiled in the presence of neutral ferrous hydroxide and titanous salts. Sodium perchlorate as well as sodium chlorate occur with sodium nitrate in native Chili saltpetre.

Composition.—10.03 grams of perchloric acid were dissolved in water, and treated with a small excess of potassium carbonate. The solution was evaporated to dryness with a slight excess of acetic acid, and washed with absolute alcohol to remove the potassium acetate. The residual potassium perchlorate was dried and weighed. The potassium perchlorate was ignited to drive off the oxygen. The results were:

Potassium perchlorate	13.8326 grams
Potassium chloride, KCl	7.4434 "
Oxygen	6.3892 "

The composition of the residual potassium chloride, by a previous analysis, is known to be KCl with a molecular weight 74.56. Divide the amount of oxygen by 16, and the amount of potassium chloride by 74.56 to get the atomic ratio KCl : O. This was found to be KCl : O = 1 : 4 corresponding with the empirical formula KClO_4 for the potassium salt, and HClO_4 for the acid. If the acid is monobasic, the formula must be HClO_4 ; if dibasic, $\text{H}_2\text{Cl}_2\text{O}_8$, etc. If the acid is dibasic, it would *probably* be possible to prepare an acid salt, KHCl_2O_8 . The acid salt has not been made, hence the analytical data may be taken as circumstantial evidence that the molecular formula of the salt is KClO_4 . Thus, it is possible to estimate the probable molecular formula of an acid by chemical analysis, and *mutatis mutandis* also of a base. By the same argument it has been shown that the formula of hydrofluoric acid is probably H_2F_2 , because it behaves as a dibasic acid.

Chlorine heptoxide.—Perchloric anhydride— Cl_2O_7 —is obtained by adding perchloric acid very slowly to phosphorus pentoxide cooled to -10° , and, after standing for a day at this temperature, distilling the mixture at 82° . The chlorine heptoxide condenses to a colourless volatile oil which decomposes in a few days. It is soluble in benzene which it slowly attacks; it also reacts with iodine, slowly in darkness, rapidly in

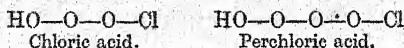
light, forming what is probably *iodine heptoxide*— I_2O_7 ; but it does not react with bromine. Chlorine heptoxide explodes in contact with a flame, and by sharp percussion. It may be poured on organic matter—paper, wood, etc.—with impunity for the oil simply volatilizes in air. It reacts with water, forming perchloric acid. Chlorine heptoxide was isolated by Michael and Conn in 1900. The preparation of the anhydride is dangerous, for, added Michael and Cohn, “the apparatus may be virtually pulverized by violent explosions, and personal precautions must be taken accordingly.”

§ 8. The Valency of Chlorine.

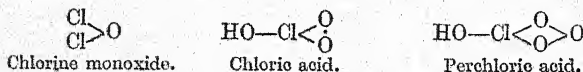
The oxides and oxyacids of chlorine so far considered are :

OXIDES.		ACIDS.	
Chlorine monoxide . . .	Cl_2O	Hypochlorous acid . . .	$HClO$
[Chlorine trioxide . . .	$Cl_2O_3]$	Chlorous acid . . .	$HClO_2$
Chlorine peroxide . . .	ClO_2
[Chlorine pentoxide . . .	$Cl_2O_5]$	Chloric acid . . .	$HClO_3$
Chlorine heptoxide . . .	Cl_2O_7	Perchloric acid . . .	$HClO_4$

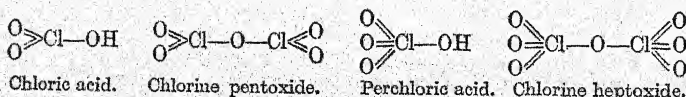
The oxides and acids show a marked increase in stability as the contained oxygen increases. The anhydrides indicated in the brackets have not been prepared, while hypochlorous and chloric acids are only known in solution. The constitution of these compounds is by no means clear. Some base the graphic formulæ on bivalent oxygen and univalent chlorine :



Carbon compounds with chains of oxygen atoms are usually less stable the longer the chain of oxygen. Here the contrary is the case, $HOCl$ is least stable, and $HOOOOCl$ is the most stable of these acids. Some, therefore, suppose that the chain formulæ are improbable, and that chlorine monoxide is constituted like nitrogen monoxide; chlorine peroxide like nitrogen peroxide, and chloric acid like nitric acid, namely—



where chlorine may be uni- and ter-valent. The nitrogen oxides will be described later. The chlorates and nitrates are isomorphous, as is also the case with the perchlorates and permanganates. This indicates a similarity in the constitution of each pair of these salts. There are therefore reasons for assuming that in chloric acid the chlorine is quinquevalent; and in perchloric acid, septivalent.



The existence of compounds like iodine trichloride, etc., is difficult to explain other than by assuming that the iodine is multi-valent, and if iodine can be multi-valent, then presumably also chlorine.

§ 9. Hypobromous and Hypoiodous Acids.

When iodine or bromine is added to a cold aqueous solution of ammonia ; or to cold solutions of potassium, sodium, calcium, or barium hydroxides, a colourless liquid is obtained which possesses bleaching qualities. This solution resembles, in many respects, corresponding solutions obtained with chlorine. Hence it is inferred that in the case of, say, iodine, hypoiodite and iodide are formed ; and in the case of bromine, hypobromite and bromide are formed. Thus with iodine : $2\text{KOH} + \text{I}_2 \rightleftharpoons \text{KI} + \text{KOI} + \text{H}_2\text{O}$. A dilute solution of the corresponding acids can be made by shaking mercuric oxide with a cold aqueous solution of bromine or iodine with water. With bromine : $\text{HgO} + \text{H}_2\text{O} + 2\text{Br}_2 = \text{HgBr}_2 + 2\text{HOBr}$. The aqueous solutions of these acids are very unstable, particularly in the case of the hypoiodites and hypoiodous acid. An aqueous solution of hypoiodous acid decomposes into hydriodic and iodic acids, and these react together forming free iodine. The aqueous solution of potassium hypoiodite decomposes at ordinary temperatures in a few hours into potassium iodide and iodate : $3\text{KOI} = 2\text{KI} + \text{KIO}_3$. Hypobromous acid decomposes at 60° into bromine and water, but the aqueous solution can be distilled at 40° in vacuo. Sodium hypochlorite or hypobromite serves as an oxidizing agent for urea and for ammonium salts. With ammonium salts the reaction is symbolized : $2\text{NH}_4\text{Cl} + 3\text{NaOBr} = 3\text{NaBr} + 2\text{HCl} + 3\text{H}_2\text{O} + \text{N}_2$. The volume of nitrogen so obtained serves as a measure of the ammonium chloride in the solution. Substances resembling bleaching powder have been formed by the action of bromine and of iodine upon slaked lime. "Iodine bleaching powder" is probably best represented, $\text{Ca}(\text{IO})\text{I}$; and "bromine bleaching powder," $\text{Ca}(\text{OBr})\text{Br}$.

Bromous acid.—What appears to be a solution of bromous acid, HBrO_2 , is formed when an excess of bromine water is agitated with a concentrated solution of silver nitrate. It is probable that hypobromous acid is first formed : $\text{AgNO}_3 + \text{Br}_2 + \text{H}_2\text{O} = \text{HOBr} + \text{AgBr} + \text{HNO}_3$; and that the hypobromous acid is subsequently oxidized : $2\text{AgNO}_3 + \text{Br}_2 + \text{H}_2\text{O} + \text{HOBr} = \text{HBrO}_2 + 2\text{HNO}_3 + 2\text{AgBr}$. However, neither the acid nor its salts have been isolated.

§ 10. Bromic Acid.

Potassium bromate— KBrO_3 .—This salt can be made by dropping 80 grams of bromine slowly into a cold solution of 62 grams of potassium hydroxide and 62 grams of water. The cold yellow solution soon deposits crystals of bromate : $6\text{KOH} + 3\text{Br}_2 = 5\text{KBr} + \text{KBrO}_3 + 3\text{H}_2\text{O}$. The crystals are filtered off, and purified by recrystallization from 130 c.c. of boiling water. The mother liquid contains potassium bromide, KBr . If an aqueous solution of barium hydroxide be substituted for potassium hydroxide, barium bromate is obtained : $\text{Ba}(\text{BrO}_3)_2 \cdot \text{H}_2\text{O}$.

Bromic acid, HBrO_3 .—An aqueous solution of bromic acid is made by treating barium bromate with the calculated quantity of sulphuric acid : $\text{Ba}(\text{BrO}_3)_2 + \text{H}_2\text{SO}_4 = \text{BaSO}_4 + 2\text{HBrO}_3$. The solution is decanted from the precipitated barium sulphate and the aqueous solution concentrated by evaporation in vacuo until it contains about 50 per cent. of the acid.

If the concentration be carried farther, or the acid be heated, it decomposes into bromine, oxygen, and water.

The bromates resemble the chlorates, but, when heated, there is no sign of the formation of perbromates analogous with perchlorates. The decomposition proceeds: $2\text{KBrO}_3 = 2\text{KBr} + 3\text{O}_2$. Some of the metallic bromates, when heated, give the metallic oxide, bromine, and oxygen: $2\text{Mg}(\text{BrO}_3)_2 = 2\text{MgO} + 2\text{Br}_2 + 5\text{O}_2$. Bromine water in aqueous solution is oxidized to bromic acid by the action of chlorine: $\text{Br}_2 + 6\text{H}_2\text{O} + 5\text{Cl}_2 = 10\text{HCl} + 2\text{HBrO}_3$. The bromic acid so obtained is mixed with hydrochloric acid. Bromic acid acts on silver without the evolution of hydrogen: $6\text{Ag} + 6\text{HBrO}_3 = 5\text{AgBrO}_3 + 3\text{H}_2\text{O} + \text{AgBr}$.

§ 11. Iodic Acid.

Potassium iodate, KIO_3 .—Potassium iodate can be made by the direct action of iodine on potassium chlorate: $2\text{KClO}_3 + \text{I}_2 = 2\text{KIO}_3 + \text{Cl}_2$. This reaction illustrates how iodine is able to replace chlorine in its oxygen compounds, although the reverse action occurs with the hydrogen compounds: $2\text{HI} + \text{Cl}_2 = 2\text{HCl} + \text{I}_2$. These differences correspond with the difference in the heats of formation (or decomposition) of the hydric acids and the oxyacids:

HYDRACIDS.		OXYACIDS.	
$\text{H} + \text{Cl} = \text{HCl} + 22.0 \text{ Cals.}$		$\text{H} + \text{Cl} + 3\text{O} + \text{Aq} = \text{HClO}_3\text{aq} + 24.0 \text{ Cals.}$	
$\text{H} + \text{Br} = \text{HBr} + 8.4 \text{ Cals.}$		$\text{H} + \text{Br} + 3\text{O} + \text{Aq} = \text{HBrO}_3\text{aq} + 12.4 \text{ Cals.}$	
$\text{H} + \text{I} = \text{HI} - 6.0 \text{ Cals.}$		$\text{H} + \text{I} + 3\text{O} + \text{Aq} = \text{HIO}_3\text{aq} + 55.8 \text{ Cals.}$	

Hence the order of the stability of the hydric acids is: HCl , HBr , HI ; and of the oxyacids: HIO_3 , HClO_3 , HBrO_3 .

To prepare potassium iodate, dissolve, say, 25 grams of potassium chlorate in a 200 c.c. flask with 120 c.c. of boiling water; add 26 grams of iodine, and 1 c.c. of concentrated nitric acid to the hot solution. In a few minutes chemical action begins and a stream of chlorine escapes from the flask. When the violence of the reaction has subsided, boil the solution to drive off the chlorine. Let the solution cool. Potassium iodate is less soluble in cold water than potassium chlorate, and accordingly most of the potassium iodate can be separated by crystallization. Filter off the crystals of potassium iodate; dissolve the crystals in hot water; neutralize the solution with potassium hydroxide; and cool the solution. Fairly pure potassium iodate crystallizes from the solution.

Barium iodate, $\text{Ba}(\text{IO}_3)_2$.—This salt can be prepared by dissolving iodine in an aqueous solution of barium hydroxide, or by the addition of barium chlorate to an aqueous solution of potassium iodate. White granular barium iodate is precipitated.

The preparation of iodic acid, HIO_3 .—Iodic acid can be made by the process described for bromic acid, of course substituting iodine for bromine, for instance, by the action of dilute sulphuric acid on barium iodate; or by the action of chlorine on water containing iodine in suspension; or by the action of an aqueous solution of chloric acid on iodine. Iodic acid, however, is usually made by the direct oxidation of iodine with nitric acid. The end products are usually represented: $10\text{HNO}_3 + \text{I}_2 = 2\text{HIO}_3 + 10\text{NO}_2 + 4\text{H}_2\text{O}$. There is, however, no doubt that the reaction is much more complex, and the equation does little more than show how iodic acid can be one product of the action of nitric acid on iodine.

Add 32 grams of powdered iodine, in small quantities at a time, to 130 grams of concentrated colourless nitric acid heated in a 500 c.c. flask fitted with a long neck to act as a condenser. A current of air, carbon dioxide, or oxygen is passed through the mixture to remove the nitrogen oxides as fast as they are formed. When the iodine has all dissolved, and a white precipitate of iodic acid has taken its place, let the solution cool. Collect the solid iodic acid on an asbestos filter. Dissolve the acid in the least possible quantity of hot water; again filter and allow the solution to recrystallize from a 20 per cent. solution of nitric acid. If the iodic acid so prepared be not colourless, it must be again crystallized, by adding to the aqueous solution an equal volume of nitric acid of sp. gr. 1.33, and concentrating the solution.

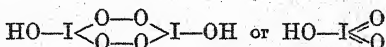
The properties of iodic acid.—Iodic acid is a white crystalline solid readily soluble in water. The aqueous solution first reddens blue litmus, and then bleaches the colour. If iodic acid be treated with concentrated sulphuric acid until iodine begins to be evolved, a yellow solid is obtained which, after washing with water, and then with ether, furnishes numbers corresponding with iodine dioxide, IO_2 , or iodine tetroxide, I_2O_4 —the molecular weight has not been determined. Iodic acid does not give a blue colour with starch. In some cases, it behaves as a monobasic acid forming salts—iodates. Most of the iodates are insoluble in water, and behave like the bromates when heated. The iodates form a series of "acid salts" with iodic acid, thus:

KIO_3
Normal iodate.

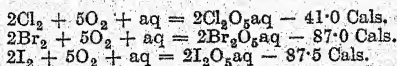
$\text{KIO}_3 \cdot \text{HIO}_3$
Uniacid iodate.

$\text{KIO}_3 \cdot 2\text{HIO}_3$
Biacid iodate.

These facts have led to the belief that, unlike the corresponding chloric acid, iodic acid is polybasic—may be $\text{H}_2\text{I}_2\text{O}_6$. Iodine may be ter- or quinque-valent in iodic acid, that is,



But really the constitution of the acid is not known with any degree of probability. When iodic acid is heated to about 170° , it decomposes with the loss of water forming iodine pentoxide, I_2O_5 , thus: $2\text{HIO}_3 = \text{H}_2\text{O} + \text{I}_2\text{O}_5$. This substance is a white crystalline solid which forms iodic acid when it is dissolved in water, hence, iodine pentoxide is iodic anhydride. A better way of making the pentoxide is to expose iodine to the action of the vapour of nitric anhydride. Both the oxidizing agent and the excess of iodine can be removed by warming the tube in which the oxidation occurred (Guichard, 1909). It is more stable than the other oxides of bromine or chlorine, but it decomposes into its elements above 200° . At this temperature the amount of decomposition in 100 hrs. is just appreciable; and at 259° the amount decomposed is proportional to the time. The above observations agree with the heats of formation of the three pentoxides:



Millon's yellow oxide I_2O_4 is probably a basic iodine iodate $\text{O} = \text{I} - \text{IO}_3$.

Iodic acid is an oxidizing agent. It oxidizes hydrogen sulphide with the liberation of iodine: $2\text{HIO}_3 + 5\text{H}_2\text{S} = 5\text{S} + 6\text{H}_2\text{O} + \text{I}_2$; and with hydriodic acid: $\text{HIO}_3 + 5\text{HI} = 3\text{I}_2 + 3\text{H}_2\text{O}$. This reaction is peculiar in that it proceeds in the presence of potassium sulphate: $5\text{KI} + \text{KIO}_3$

+ $3\text{H}_2\text{SO}_4 = 3\text{K}_2\text{SO}_4 + 3\text{H}_2\text{O} + 3\text{I}_2$, more slowly at 80° than at 18° . In the very great majority of cases so far noticed reactions progress faster the higher the temperature. Iodic acid with sulphur dioxide: $2\text{HIO}_3 + 5\text{SO}_2 + 4\text{H}_2\text{O} = 5\text{H}_2\text{SO}_4 + \text{I}_2$. This latter reaction is used as a test for iodates. The solution is first acidified with hydrochloric acid to liberate the iodic acid, and then mixed with starch paste. Sulphurous acid, or an alkaline sulphite, is then added drop by drop. The liberated iodine forms "blue starch iodide," the characteristic reaction of iodine.

The period of induction.—The reaction between sulphurous acid and an iodate is very interesting, because the iodine does not appear immediately the substances are mixed; there is a well-defined period of time—period of induction—between the moment the reacting substances are mixed, and the moment iodine makes its appearance. The duration of the interval is dependent upon the concentrations of the solutions. This can be demonstrated in the following manner: Dissolve 1.8 gram of iodic acid in a litre of water; also prepare a litre of an aqueous solution of 0.9 gram of

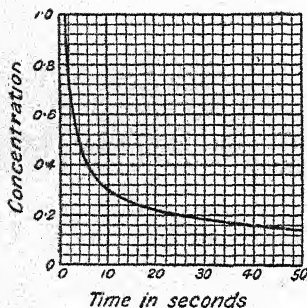


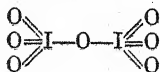
FIG. 132.—Period of Induction.

of sodium sulphite, $\text{Na}_2\text{SO}_3 \cdot 7\text{H}_2\text{O}$; 10 per cent. sulphuric acid, and 9.5 grams of starch made into a paste with hot water. Add 100 c.c. of each solution to separate beakers; and mix the two. Note the time when the solutions are mixed. Count the seconds which pass before the starch blue appears. Dilute each solution to 0.8, 0.6, 0.4, 0.2, of its former concentration, and repeat the experiments with the diluted solutions. Plot the results as has been done in Fig. 132. If the concentration and temperature be constant, the same results can always be reproduced. It is supposed that the first action is due to the reduction of the iodic acid to hydriodic acid: $3\text{H}_2\text{SO}_3 + \text{HIO}_3 = 3\text{H}_2\text{SO}_4 + \text{HI}$; and when all the sulphurous acid has been oxidized, this reaction is followed by the reduction of the hydriodic acid by the excess of iodic acid: $\text{HIO}_3 + 5\text{HI} = 3\text{H}_2\text{O} + 3\text{I}_2$. The first reaction—oxidation of sulphurous acid—must be nearly completed before the second one can start, because, as indicated above, the iodine with sulphurous acid reforms hydriodic acid. Since the maximum amount of iodic and sulphurous acids are present at the start, the first-named reaction must be fastest at the beginning and afterwards gradually slow down. Again, since the velocity of a reaction is proportional to the amount of the reacting substances present in the system, the second reaction will be the slowest at the start, and gradually become faster. The speed of formation of the iodine is therefore the resultant velocity of two consecutive reactions; and the belated appearance of the iodine—the period of induction—corresponds with the time required for the first reaction to make enough hydriodic acid to enable the second reaction to make sufficient iodine to colour the starch. To summarize the successive steps in the reduction of iodic acid to iodine: $2\text{HIO}_3 \rightarrow 2\text{HI} \rightarrow \text{I}_2$, where HI

represents an "intermediate compound" characteristic of consecutive reactions generally.

§ 12. Periodic Acid, and Perbromic Acid.

It was once thought that perbromic acid— HBrO_4 —could be made by the action of bromine on perchloric acid; on perchlorates in the presence of potassium bromide; by the action of bromine on sodium periodate; or by the action of potassium bromate on lead dioxide; but the evidence is not satisfactory. What was taken to be perbromic acid was really a mixture of perchloric and sulphuric acids, and it is therefore somewhat doubtful if perbromic acid has yet been prepared. It is not certain if normal periodic acid— HIO_4 —has been made, but a large number of related salts, classed as periodates, are known. These are somewhat difficult to understand because of their complexity. Our study of perchloric acid may help us a little. Starting from the assumption that iodine is septivalent in periodic anhydride— I_2O_7 ; the constitutional formula of iodine heptoxide will be:



This compound has not yet been prepared; when it is, we infer that it will be analogous with the known corresponding compound of chlorine— Cl_2O_7 . By adding water step by step to this oxide, we can imagine the series of acids with the corresponding salts represented in the following table. Many salts have been isolated even when the corresponding acids are unknown.

TABLE XIV.—THE PERIODIC ACIDS.

I_2O_7 with	Compounds formed.	Empirical formula of acid.	Name of acid.	Salts.
$1\text{H}_2\text{O}$	$\text{I}_2\text{O}_6(\text{OH})_2$; or $2\text{IO}_3\text{OH}$	HIO_4	Meta-periodic acid	KIO_4 ; AgIO_4 ; etc.
$2\text{H}_2\text{O}$	$\text{I}_2\text{O}_5(\text{OH})_4$	$\text{H}_4\text{I}_2\text{O}_9$	Dimeso-periodic acid	No salts known
$3\text{H}_2\text{O}$	$\text{I}_2\text{O}_4(\text{OH})_6$; or $2\text{IO}_2(\text{OH})_3$	H_3IO_5	Meso-periodic acid	Ag_3IO_5 ; $\text{Pb}_3(\text{IO}_5)_2$; etc.
$4\text{H}_2\text{O}$	$\text{I}_2\text{O}_3(\text{OH})_8$	$\text{H}_8\text{I}_2\text{O}_{11}$	Dipara-periodic acid	Salts have been reported; prob- ably mixtures
$5\text{H}_2\text{O}$	$\text{I}_2\text{O}_2(\text{OH})_{10}$; or $2\text{IO}(\text{OH})_5$	H_5IO_6	Para-periodic acid	Ag_5IO_6 ; $\text{Ag}_5\text{H}_5\text{IO}_6$; etc.
$6\text{H}_2\text{O}$	$\text{I}_2\text{O}(\text{OH})_{12}$	$\text{H}_{12}\text{I}_2\text{O}_{13}$	Diortho-periodic acid	Salts have been reported; prob- ably mixtures
$7\text{H}_2\text{O}$	$\text{I}_2(\text{OH})_{14}$; or $2\text{I}(\text{OH})_7$	H_7IO_7	Ortho-periodic acid	No salts known

Nomenclature of the periodic acids and the periodates.—When one anhydride forms a series of acids by union with different amounts of water, the acid containing most water is called the ortho-acid—from the Greek *ὀρθός* (orthos), regular; the other acids have prefixes, making para-acid—from the Greek *παρά* (para), beside; meso-acid—from the Greek *μέρος* (meros), middle, intermediate;

and meta-acid—from the Greek *μετά* (meta), beyond, less than the highest. The di-acids are supposed to be formed by abstraction of one molecule of water from two molecules of acid. It will be evident that if the acids are polybasic, we can imagine the available hydrogen atoms replaced one by one. If all the available hydrogen atoms are replaced by bases, the normal salt is obtained; if only one of the available hydrogen atoms is replaced, the primary salt; if two, the secondary salt; if three, the tertiary salt.

Secondary sodium para-periodate has the composition $\text{Na}_2\text{H}_3\text{IO}_6$, thus corresponding with para-periodic acid, or $\text{HIO}_4 \cdot 4\text{H}_2\text{O}$. It can be prepared as a sparingly soluble salt by passing chlorine into an aqueous solution of sodium iodate and sodium hydroxide: $\text{NaIO}_3 + 3\text{NaOH} + \text{Cl}_2 = 2\text{NaCl} + \text{Na}_2\text{H}_3\text{IO}_6$. When sodium periodate is dissolved in nitric acid and silver nitrate is added to the solution, crystals of silver meta-periodate— AgIO_4 —corresponding with meta-periodic acid, separate when the solution is concentrated by evaporation.

Para-periodic acid.—When silver meta-periodate is boiled with water, an insoluble salt, secondary silver para-periodate, is formed along with para-periodic acid, H_5IO_6 , or $\text{HIO}_4 \cdot 2\text{H}_2\text{O}$. Thus $2\text{AgIO}_4 + 4\text{H}_2\text{O} = \text{Ag}_2\text{H}_3\text{IO}_6 + \text{H}_5\text{IO}_6$. When the clear solution is concentrated, deliquescent crystals of the para-periodic acid are obtained. This acid melts at 133° , and decomposes into iodine pentoxide, water, and oxygen at 150° . The water cannot be expelled by heat because oxygen is evolved along with the water. Periodic acid and neutral alkali periodates are reduced to iodates by hydrogen peroxide. If concentrated solutions of periodic acid and an excess of hydrogen peroxide be used, much iodine is liberated at the same time as the oxygen. Iodic acid and concentrated solutions of sodium iodate are slowly reduced by 3 per cent. hydrogen peroxide.

A small quantity of potassium periodate is said to occur in native Chili saltpetre, along with some sodium iodate.

§ 13. The Halogen Family of Elements.

Fluorine, chlorine, bromine, and iodine together form a remarkable family of elements. Following J. J. Berzelius, the whole group is sometimes called the halogens—from the Greek *ἅλς* (hals), sea-salt; *γεννάω* (gennao), I produce. This name was applied because these elements—principally chlorine, bromine, and iodine—are found in sea-water, and the sodium salts resemble sea-salt (sodium chloride). The fluorides, chlorides, bromides, and iodides are often called the halides. The gradation in characters which the halogens show with increasing atomic weights from fluorine to iodine yield an almost perfect family series. The family relationship of the halogens is illustrated by—

(1) The *similarity* in the chemical and physical properties of the elements and their corresponding compounds is such that the properties of any one member of the family can be said to summarize or typify the properties of the other members of the series.

(2) The *gradual transition* of chemical and physical properties such that if the elements be arranged in order: F, Cl, Br, I, the variation in any particular property in passing from fluorine to iodine nearly always proceeds in the same order, and that is the order of their atomic weights.

The relationship in the physical and chemical properties of the halogens can best be emphasized by the tabular scheme.

TABLE XV.—SOME PROPERTIES OF THE HALOGENS.

Property.	Fluorine.	Chlorine.	Bromine.	Iodine.
Atomic weight	19	35.46	79.92	126.29
State of aggregation . .	Gas	Gas	Liquid	Solid
Melting point	-223°	-102°	-7.3°	+113°
Boiling point	-187°	-33.6°	58° to 63°	183°
Specific gravity	1.14 (liq.)	1.55 (liq.)	3.19 (liq.)	5 (solid)
Atomic volume ¹	16.7	22.9	25.1	25.6
Colour of gas	Pale yellow	Greenish yellow	Brownish red	Violet
Solubility (100 grms. water at 20°) dissolves grams.	Decomposes water	0.716	3.2	0.03
Oxidizing action	Strongest	Very strong	Strong	Weak
Oxyacids	—	HClO	HBrO	HIO
	—	HClO ₂	—	—
	—	HClO ₃	HBrO ₃	HIO ₃
	—	HClO ₄	—	H ₅ IO ₆
Halides	HF	HCl	HBr	HI

All the halogens form compounds with hydrogen, and the readiness with which union occurs decreases as the atomic weight increases. The properties of the halogen acids and their salts show as striking a relationship as the elements themselves, and are shown in Table XVI.

TABLE XVI.—THE PROPERTIES OF THE HALOID ACIDS.

Property.	Hydrogen fluoride.	Hydrogen chloride.	Hydrogen bromide.	Hydrogen iodide.
Molecular weight . . .	20	36.46	80.93	127.93
Boiling point	19.4°	-83.4°	-67.1°	-35.5 (4 atm.)
Melting point	-92.3°	-113°	-87.9°	-50.8°
Solubility in water . .	35.3 %	42 %	49 %	57 %
Specific gravity saturated aq. sol.	1.15	1.21	1.49	1.70
Boiling point aq. sol. .	111° (43 %)	110° (20 %)	126° (47 %)	127° (57 %)
Dissociates at	—	1500°	800°	180°
Heat of formation (Cals.)	+ 38.5 gas	+ 22.0 gas	+ 12.3 gas	+ 0.4 gas
Heat of formation of K salt (Cals.)	110.6	105.7	95.3	80.1
Heat of neutralization NaOH	68 Cals.	57 Cals.	57 Cals.	57 Cals.
Potassium salt melts . .	885°	790°	750°	705°
Calcium salt melts . . .	1330°	780°	780°	740°
Solubility Ag salt (20°) per 100 grms. water . .	1720	0.00154	0.000084	0.0000028
Solubility Ca salt per 100 c.c. solution	0.0016	42.7	58.8	67.6

Taking almost any property and comparing its magnitude in passing from fluorine to iodine, or from fluorides to iodides, a similar gradual

¹ Atomic volume is the atomic weight divided by the specific gravity.

transition will be observed. Thus, with the melting points of some halides :

	Fluoride.	Chloride.	Bromide.	Iodide.
Calcium	1361°	780°	780°	740°
Barium	1280°	960°	880°	740°
Sodium	980°	820°	765°	650°
Potassium	885°	790°	750°	705°

The jump in passing from the fluorides to the chlorides is greater than with the other steps in the series. With the alkaline earths this is supposed to be connected with the dibasicity of hydrofluoric acid. The solid elements form isomorphous rhombic crystals; and the halide salts are also usually isomorphous.

The boiling point and composition of the acid of constant composition of hydrofluoric acid is exceptional. Hydrogen fluoride, HF, appears to be polymerized below 88°. Again, while the affinity of the halogens for hydrogen decreases with increasing atomic weight, the reverse is the case with oxygen. But with oxygen the relationship is not so clearly defined. Thus although fluorine forms no known compound with oxygen, numerous compounds of oxygen with chlorine have been obtained; and, judging by the known compounds with oxygen, the affinity of bromine for oxygen appears to be less, not greater, than is the case with chlorine, although we must remember that bromine has not been investigated so much as chlorine. Bromine occupies an anomalous position with respect to oxygen. Chlorine, bromine, and iodine unite with one another forming an unstable series of compounds analogous with hypochlorites and chlorates, but the bromine analogue of perchlorates has not been prepared.

Questions.

1. You are required to establish experimentally the validity or otherwise of the following equation: $6\text{KOH} + 3\text{Cl}_2 = \text{KClO}_3 + 5\text{KCl} + 3\text{H}_2\text{O}$. Describe in detail the methods you would employ.—*Science and Art Dept.*

2. Describe the preparation and properties of the oxides and oxy-acids of chlorine. What is understood by an *endothermic compound*, and by what peculiarities are such compounds distinguished?—*Aberdeen Univ.*

3. Make out a table comparing the various members of the chlorine family.—*Princeton Univ., U.S.A.*

4. Calculate the percentage of "active chlorine" in a sample of bleaching powder from the following data: 10 grams of the powder was extracted with 500 c.c.m. of water; 25 c.c.m. of the extract when added to an acidified solution of potassium iodide set free iodine equivalent to 32.7 c.c.m. of $\frac{1}{10}$ N solution of sodium thiosulphate ($\text{Cl} = 35.5$; $\text{I} = 127$).—*Univ. North Wales.*

5. Describe the methods by which iodic and periodic acids are best prepared, and discuss the evidence afforded by these compounds as to the valency of iodine.—*Science and Art Dept.*

6. How can chlorine be prepared from bleaching powder? What is the effect of passing chlorine into an aqueous solution of (a) hydrogen sulphide, (b) sulphur dioxide, (c) potassium hydroxide, (d) potassium iodide?—*Sheffield Univ.*

7. Point out the chemical similarity between chlorine, bromine and iodine. What is this group called? How is each of these elements set free from a salt containing it?—*Sheffield Scientific School, U.S.A.*

8. Three salts are known which contain oxygen in addition to potassium and chlorine. Write the formulae of these salts and explain how each salt can be (a) prepared, (b) recognized, (c) made to furnish chlorine.—*Sheffield Univ.*

9. Make a chart of the halogens (a) in the order of the atomic weights; (b) giving the colour of each; (c) giving the physical state of each; (d) and the relative stability of the common acids.—*Amherst Coll., U.S.A.*

10. Give the law of multiple proportion, illustrating from the series of chlorine oxyacids.—*Amherst Coll., U.S.A.*

11. It has long been known that the decomposition of potassium chlorate is effected at a lower temperature in the presence of certain metallic oxides, that bleaching powder will yield oxygen in the presence of cobalt oxide, and that oxygen is evolved from hydrogen peroxide by the addition of finely divided silver or platinum. What explanations can you give of such changes?—*Science and Art Dept.*

12. Compare and contrast the properties of hydrogen fluoride and of hydrogen chloride; and the properties of the perchlorates with those of the periodates.—*Owens Coll.*

13. "The most important compounds of chlorine with hydrogen and oxygen are hypochlorous acid, chloric acid, and perchloric acid." Describe how you would prepare each of these acids from chlorine as the starting product. By what reactions could you distinguish hydrochloric acid from hypochlorous acid?—*London Univ.*

14. How do you account for the greater stability of perchloric as compared with chloric acid?—*Board of Educ.*

15. What is meant by saying that fluorine is a "halogen"? Compare its properties with those of the other halogens, and describe the preparation of two fluorides, one of a metal, and the other of a non-metal, beginning with the fluorine in the form of fluorspar.—*London Univ.*

16. Trace the successive reactions involved in the production of hypochlorites, chlorates, and perchlorates, discussing for each stage the chemical change which occurs in it, and the conditions under which it takes place. By what properties are these salts severally distinguished?—*Board of Educ.*

17. What products may be obtained by the electrolysis of sodium chloride under different conditions? State the conditions in each case, and give equations when possible.—*Panjab Univ.*

18. Calculate the volume of moist nitrogen at 20° and 770 mm. that would be theoretically obtained by the action of 0.25 gram. of ammonium chloride upon a solution of sodium hypochlorite ($H = 1$, $N = 14$, $Cl = 35.5$; pressure of aqueous vapour at $20^{\circ} = 17.4$ mm.).—*Madras Univ.*

19. Describe the process of bleaching with bleaching powder. What precautions have to be taken in preserving bleaching powder? Give reasons. How would you compare in an exact manner two specimens of bleaching powder?—*Calcutta Univ.*

20. How is perchloric acid prepared and how is it distinguished from hypochlorous and chloric acids.—*Science and Art Dept.*

21. Describe the methods by which iodic and periodic acids are best prepared, and discuss the evidence afforded by these compounds as to the valency of iodine.—*Science and Art Dept.*

22. How can hydriodic acid and iodic acid be severally prepared from iodine? What is the action of silver nitrate on each in aqueous solution, and what happens when aqueous solutions of these acids are mixed?—*London Univ.*

23. Write an account of iodic acid and the iodates.—*Science and Art Dept.*

24. 20 grms. of bleaching powder were made up with water to a litre, and 25 c.c. were titrated with a standard sodium arsenite of such a strength that 1 c.c. corresponded with 0.00355 gram. of chlorine. After adding 40 c.c. of the standard solution, a drop of the liquid being titrated gave no coloration with starch and potassium iodide. Required the percentage amount of available chlorine. Here $46 \times 0.00355 = 0.163$ gram. of chlorine per 25 c.c. of solution, or 6.5 grms. per 1000 c.c. of solution or 20 grms. of sample, etc.

CHAPTER XVI

ELECTROLYSIS AND THE IONIC HYPOTHESIS

§ 1. The Products of Electrolysis.

The electricity which decomposes and that which is evolved by the decomposition of a certain quantity of matter are the same.—M. FARADAY.

ONE or both of the products of electrolysis may be an insoluble solid, a soluble liquid, a gas, etc. When an insoluble solid is formed, it will either stick to the electrode or fall to the bottom of the electrolytic cell (Fig. 7); if a gas, not too soluble in the electrolyte, be formed, the gas can be collected as indicated in Fig. 19. Soluble substances are not always visible when in solution. The soluble matters can often be isolated more or less completely by surrounding the proper electrode with a porous pot which retards the diffusion and mixing of the products separated at the two electrodes. This is done, for example, in the industrial preparation of chlorine.

It will be remembered that during the electrolysis of copper sulphate, the products of the electrolysis were: copper, sulphuric acid, and oxygen. This is more than was present in the copper sulphate used at the start. It is therefore assumed, as a trial hypothesis, that ions Cu and SO_4 are produced at the electrodes during the passage of the current; that the copper cation carries a positive charge of electricity, and the SO_4 anion a negative charge. Consequently, the Cu ion will be found at the negative electrode, and the SO_4 ion at the positive electrode. The ions are de-electrified at the electrodes: the Cu at the cathode, and the SO_4 at the anode. The de-electrified copper ions are deposited as metallic copper about the cathode; and the de-electrified SO_4 , at the anode, reacts at once with the solvent (water), producing sulphuric acid and oxygen: $2\text{SO}_4 + 2\text{H}_2\text{O} = 2\text{H}_2\text{SO}_4 + \text{O}_2$. When an aqueous solution of potassium nitrate is electrolyzed, potassium hydroxide and gaseous hydrogen are formed at the cathode; and nitric acid and oxygen at the anode. It is assumed that the potassium nitrate is first decomposed into two electrified ions, K and NO_3 , at the electrodes; and that the K ion, when de-electrified, reacts with water at the cathode, producing potassium hydroxide and hydrogen; and the NO_3 ion, when de-electrified at the anode, reacts with water, giving nitric acid and oxygen: $4\text{NO}_3 + 2\text{H}_2\text{O} = 4\text{HNO}_3 + \text{O}_2$.

Again, if a solution of silver nitrate be electrolyzed with silver electrodes, metallic silver is deposited at the cathode, and the nitric acid produced at the anode attacks and dissolves the cathode-forming silver nitrate. This explains how the total concentration of a solution of silver nitrate does not alter if electrolyzed in a cell with a silver anode. The nitric acid

produces more silver nitrate as fast as it is formed. Similar remarks apply to the electrolysis of solutions of copper sulphate with a copper anode, etc.; of ferrous ammonium sulphate with an iron anode; of nickel ammonium sulphate with a nickel anode; etc.

Electroplating.—If copper sulphate were used as electrolyte in place of zinc sulphate, Fig. 7, metallic copper would be precipitated at the cathode, and sulphuric acid would have been formed about the anode; with silver nitrate, metallic silver would collect about the cathode, and nitric acid about the anode. If a plate of silver be used as the anode during the electrolysis of silver nitrate, metallic silver will be dissolved by the nitric acid as fast as the acid is formed. Thus, the concentration of the silver nitrate in the solution will remain unchanged and metallic silver will be transported to the cathode. This is the principle of the method of electroplating. In the case of *silver plating* a firmer and more uniform deposit of silver is obtained by using a solution of silver cyanide in potassium cyanide as the electrolyte in place of silver nitrate. The article to be plated, say a brass spoon, is attached to a wire and dipped in the solution of silver salt, and this is made the cathode. A bar or sheet of silver is made the anode. A rather weak electric current is sent through the electrolyte. The electrolyte is decomposed, and silver (cation) is deposited on the article to be plated (cathode); the anion collecting at the anode dissolves the silver anode, and thus keeps the strength of the electrolyte unchanged. What is dissolved at the anode is deposited at the cathode. Salts of other metals—nickel, iron, gold, copper, platinum, etc.—can be used as electrolytes in a similar manner, and accordingly articles can be nickel-plated, gold-plated, etc. The plated articles may be afterwards burnished.

We first inquire if there is any relation between the quantity of electricity passing through an electrolytic cell and the amount of decomposition. In order to fix a standard of measurement, let the quantity of electricity required to deposit 0.001118 gram of silver be called a coulomb. This is the so-called unit quantity of electricity. Hence 108 grams of silver, that is, a chemical equivalent of silver, will be deposited by 96,540 coulombs of electricity. This amount of electricity is often called a farad. The quantity of an element liberated by the passage of one farad of electricity is called the electro-chemical equivalent of the element.

The so-called "hydraulic analogy" of an electric current might here be cited to remind the student what he probably learned in "Physics." The quantity of water flowing through a pipe can be expressed in gallons or cubic feet per second; in a similar way, quantity of electricity may be expressed in terms of coulombs per second. An electric current carrying one coulomb per second is called an ampere. This is the so-called unit current of electricity—e.g. if 60 coulombs of electricity pass through a circuit in 20 seconds, the current is $60 \div 20 = 3$ amperes (i.e. 3 coulombs per second). A coulomb by the analogy would correspond with, say, a gallon or a cubic foot of water; and an ampere with a gallon or cubic foot of water per second. The total quantity of water delivered by a pipe is determined by the "head" or pressure of water, so that in order to pass a certain number of gallons per second through a given pipe, a certain pressure must be applied to overcome the frictional resistance of the pipe. In the same way, a certain electromotive force—electrical pressure—is required on account of the resistance offered by the wire to the flow of electricity. Just as water pressure is measured in pounds per square inch, or in feet "difference of level" or "head," so the unit of electrical pressure, the volt, is the difference of potential needed to produce a current of one ampere in a conductor whose resistance is equivalent to that of a uniform column of 14.45 grams of mercury, 106.3 cm. long. The resistance of such a column is called an ohm. Hence a volt is the electric pressure

required to produce a current of one ampère in a conductor of one ohm resistance. The terms **voltage**, **electrical pressure**, and **electromotive force** are generally applied synonymously to an electric current, or, if the current is not directly under consideration, the term **difference of potential** is used. It is of course needless to dwell on the fact that the analogy used above in comparing an electric current with a moving fluid is merely a convenience. It is probable that electricity is not a fluid, and the analogy must not be carried much further.

§ 2. Faraday's Laws of Definite Electrolytic Action.

Nature presents us with a single definite quantity of electricity. For each chemical bond which is ruptured within an electrolyte, a certain quantity of electricity traverses the electrolyte which is the same in all cases.—G. J. STONEY.

M. Faraday (1834) found that the amount of chemical work done by an electric current is directly proportional to the quantity of electricity which passes through the electrolyte. If one farad leads to the separation of 108 grams of silver, two farads will lead to the separation of 216 grams of silver, and so on. Similar results are obtained with other electrolytes. Hence, said Faraday, "the chemical decomposing action of a current is constant for a given quantity of electricity." Or "the quantity of chemical decomposition is exactly proportionate to the quantity of electricity which has passed through the electrolyte"; and consequently,

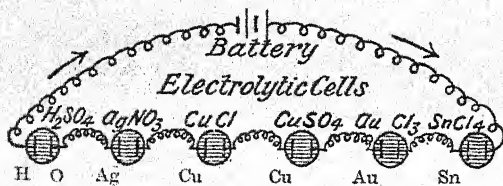


FIG. 133.—Experiment illustrating Faraday's Laws.

"the products of decomposition . . . afford a very excellent and valuable measure of the electricity concerned in their evolution." The increase in the weight of, say, the

negative electrode during the electrolysis of silver nitrate, or copper sulphate owing to the deposition of metallic silver or copper respectively, is a measure of the quantity of electricity which has passed through the system. A cell specially designed for such measurements is called a silver voltameter or a copper voltameter respectively.

Provided there are no disturbing secondary actions, the amount of electro-decomposition is not affected by the strength (or intensity) of the current, the time the current is passing, the concentration of the solution, the nature of the dissolved substance, nor by the temperature. The same quantity of electricity will always liberate the same quantity of the elements stated. The accuracy of the law is said to have been established for "currents so small that a century would be required for the separation of a milligram of hydrogen" and in large electrochemical works, the law is continually being verified by the passage of millions of coulombs. In every case, the law describes the phenomena exactly.

Again, let a current be *simultaneously* passed through six cells containing respectively dilute sulphuric acid, aqueous solutions of silver nitrate, cuprous chloride, cupric sulphate, gold chloride, and stannic chloride. The experiment is conducted by arranging the electrolytic cells as illustrated in the plan, Fig. 133. After about half an hour's electrolysis the amounts

of the different elements collected at the cathode can be weighed or measured. The results will be very nearly :

	Dilute H_2SO_4		AgNO_3	CuCl (acid sol.)	CuSO_4	AuCl_3	SnCl_4
	Cathode.	Anode.					
	Hydrogen.	Oxygen.	Silver.	Copper.	Copper.	Gold.	Tin.
Amount found	0.266	0.2126	2.9370	1.6900	0.8440	1.7476	0.7554 gram
If $\text{H} = 1$	1	8	108	63.5	31.8	65.7	29.8
Atomic weight	1.01	16	107.9	63.6	63.6	197.2	119
Valency	1	2	1	1	2	3	4

Accordingly, chemically equivalent quantities of the different elements (that is, atomic weight \div valency) are liberated by the passage of the same quantity of electricity. Consequently, the electrochemical equivalent of an element is numerically the same as the chemical equivalent.

At first sight, this result appears to contradict the principle of excluded perpetual motion, because, if the current from a $\text{Zn}|\text{H}_2\text{SO}_4|\text{Pt}$ battery be sent through an indefinite number of electrolytic cells containing dilute sulphuric acid, the same amount of hydrogen would be liberated in each, and sufficient hydrogen could be collected to furnish, on combustion, enough heat to evaporate the solution of zinc sulphate in the battery to dryness, to transform the zinc sulphate to metallic zinc and sulphuric acid, and so reconstruct the battery; and have some hydrogen remaining in excess. The experiment fails. The current will not traverse an indefinitely large number of cells. W. H. Wollaston showed in 1801 that in dealing with electrical energy we are concerned with two different factors, and that "quantity of electricity" is only one of these factors. Faraday's law describes the influence of "quantity of electricity" upon electrolysis. It says nothing about the electrical pressure—the electromotive force, described in volts—required to drive the necessary quantity of electricity through the system. Hence Faraday's work may be summarized: The same quantity of electricity passing through one or more electrolytes connected up in series, will liberate in each cell chemically equivalent amounts of the products of electrolysis, provided the electromotive force permit the necessary current to be maintained. It might here be added that, for reasons which will be discussed later, a certain specific electrical pressure or voltage—called the decomposition voltage—is required to electrolyze a given solution; thus, hydrochloric acid requires about $1\frac{1}{2}$ volts, and fused sodium chloride about 4 volts.

Let each positive charge of electricity be represented by a small dot, and each negative charge by a small dash at the upper right-hand corner of the chemical symbol for an element, then, a silver ion will be written Ag^{\cdot} ; a zinc ion by $\text{Zn}^{\cdot\cdot}$; a nitrate ion, NO_3^{\cdot} ; and a sulphate ion by $\text{SO}_4^{\cdot\cdot}$. In the electrolysis of aqueous solutions of salts, etc., the separation of an ion at one electrode is always attended by the separation of a chemically equivalent ion or ions at the other electrode. For instance, with zinc chloride, for every $\text{Zn}^{\cdot\cdot}$ which is de-electrified at the cathode, two Cl^{\cdot} ions will be de-electrified at the anode.

We have just seen that the electrochemical and chemical equivalents are numerically the same, and therefore the electrochemical equivalent of an element is obtained by dividing the atomic weight by the valency. The same quantity of electricity—positive or negative—must therefore be carried by each univalent atom, and accompany it in all its movements in the electrolytic fluid. This quantity has been called unit charge of

the ion. This seems rather startling, for it seems to imply that an electric charge is divisible not into fractions of a charge, but there may be n -multiples of unit charge so that a univalent ion carries one charge of electricity (96,540 coulombs), a bivalent ion, two charges, and an n -valent ion, n charges. According to this view, valency represents the number of charges of electricity which are associated with the respective ions, and chemically equivalent quantities of matter have the same capacity for electricity. The chemical equivalent is the electrical unit of matter; or, as Faraday expressed it:

The equivalent weights of bodies are simply those quantities which contain equal quantities of electricity, or have naturally equal electric powers; it being electricity which determines the equivalent number, because it determines the combining force. Or, if we adopt the atomic theory or phraseology, then the atoms of bodies which are equivalent to each other in their ordinary chemical action have equal quantities of electricity naturally associated with them.

Quantity of electrolytic work done by a current.—M. Faraday has shown that the amount of chemical decomposition in a given time depends upon the amount of current employed; and that the one magnitude can be computed when the other is known. To find the relation between the chemical equivalent of an element and quantity of electricity, let ϵ denote the chemical equivalent of a substance; then the weight w of an element liberated by constant quantity of electricity is proportional to ϵ ; but w is also proportional to the quantity of electricity C ; and therefore $w = k\epsilon C$, where k is a constant sometimes called *Faraday's constant*. If suitable units be chosen, $w = \epsilon C$. Consequently, the electrochemical equivalent of an element is the amount in grams liberated by one coulomb. Careful measurements have shown that 0.001118 grm. of silver will be deposited by one coulomb. Since the chemical equivalent of silver is very nearly 107.88, hydrogen unity, 0.001118 grm. is 107.88 times greater than the amount of hydrogen which will be separated by a coulomb; accordingly, 0.0001036 grm. or 1.036×10^{-5} grm. of hydrogen will be separated by the passage of one coulomb of electricity. The electrochemical equivalent of univalent copper is $63.5 \times 1.036 \times 10^{-5}$; of bivalent copper, $\frac{1}{2}$ of $63.5 \times 1.036 \times 10^{-5}$; of ferrous iron, $\frac{1}{2}$ of $56 \times 1.036 \times 10^{-5}$; of ferric iron, $\frac{1}{3}$ of $56 \times 1.036 \times 10^{-5}$; and generally, if ϵ denotes the chemical equivalent of an element—that is, atomic weight \div valency—the electrochemical equivalent is $1.036 \times 10^{-5}\epsilon$. Again, if a coulomb of electricity liberates 1.036×10^{-5} grms. of hydrogen per second, a ampères will liberate $1.036 \times 10^{-5}a\epsilon$ grms. per second, and generally, *the number of grams of an element whose chemical equivalent is liberated by the passage of a ampères of electricity flowing for t seconds is $1.036 \times 10^{-5}\epsilon at$* . If ϵ denotes the electrochemical equivalent of an element, such that $\epsilon = 1.036 \times 10^{-5}\epsilon$, the number of grams of an element whose electrochemical equivalent is ϵ liberated by the passage of a ampères flowing for t seconds is ϵat . It also follows that $1.036 \times 10^{-5}A/n$ grms. of an n -valent element of atomic weight A , will be deposited per coulomb. By Avogadro's rule, a gram-molecule of a gas—of molecular weight M and N -atoms per molecule—occupies 22.4×10^3 c.c. Hence $(22.4 \times 1.036 \times 10^{-5})/Nn$ c.c. of the gas are given off per coulomb. For a univalent gas with two atoms per molecule, $N = 1$, $n = 2$, and therefore 0.116 c.c. are obtained per coulomb at n.p.t., or 8.6193 coulombs are required per c.c. of gas. This

corresponds with 0.4176 litre per ampère hour, or 2.3943 ampères are needed per litre of gas; or with 14.750 c. ft. are liberated per 100 amp. hours, or 67.798 amps. are needed per cubic foot of gas.

EXAMPLES.—(1) An electric current is passed simultaneously through the following solutions: Hydrochloric acid, ferrous sulphate, ferric sulphate, and silver potassium cyanide. If 5.2 litres of hydrogen at n.p.t. were evolved from the hydrochloric acid, how much metal would be deposited in the case of the iron and silver salts? Here 5.2 litres of hydrogen weigh 0.4664 gm. The chemical equivalent—that is, one gram—of hydrogen is equivalent to $\frac{1}{2}$ of 56 = 28 grms. of ferrous iron; to 18.67 grms. of ferric iron; and to 108 grms. of silver. Hence $0.4664 \times 28 = 13$ grms. of iron will be deposited from the ferrous sulphate; 8.7 grms. from the ferric sulphate; and 50.1 grms. from the silver solution.

(2) A current of 0.04 amp. was passed through a solution of copper sulphate for $1\frac{1}{2}$ hrs. Hence, how much copper was deposited when the electrochemical equivalent of copper is 0.000329 gm. ? There are 5400 seconds in $1\frac{1}{2}$ hrs., hence $0.000329 \times 0.04 \times 5400$ grms. of copper were deposited.

(3) A current of $2\frac{1}{2}$ amps. was obtained from a voltaic cell for $2\frac{1}{2}$ hrs. How much zinc was dissolved, given the electrochemical equivalent of zinc is 0.000337 ? Ansr. 6.8 grms.

(4) What are the electrochemical equivalents ϵ of hydrogen, of copper (cupric), and of zinc? Hydrogen, $1.036 \times 10^{-5} \times 1 = 0.0001036$; copper, $\frac{1}{2}$ of $63.5 \times 1.036 \times 10^{-5} = 0.000329$; zinc, $\frac{1}{2}$ of $65 \times 1.036 \times 10^{-5} = 0.000337$.

(5) In C. Höpfner's method (1900), a solution of cupric and sodium chlorides is used as electrolyte and the copper passes into solution as cuprous chloride. Compare the quantity of electrical energy required to precipitate copper from Höpfner's solution, and from a solution containing copper sulphate. Here bivalent copper requires $2 \times 96,540$ coulombs for precipitating 63.57 grms., while univalent copper requires but 96,540 coulombs for precipitating the same weight of copper. Hence the precipitation of a given weight of copper from cuprous chloride requires but half the electrical energy as that required when cupric sulphate is used.

H. von Helmholtz (1881) has emphasized the fact that the evidence indicates that electricity associates with the atoms of matter in multiples of one fundamental quantity, but never in fractions of it; these fractions may not be impossible, but they have not yet been found. Hence the evidence for the atomic nature of electricity is much the same as for the atomic nature of matter. The charge on a monad atom is therefore a natural unit of electricity. To illustrate the prodigious electrical capacity of the molecules, H. von Helmholtz estimates that if the opposite electricities were extracted from a milligram of water, and given to two spheres a mile apart, these two spheres would attract each other with a force of ten tons.

Berzelius' dualistic theory.—According to J. J. Berzelius (1812) each element possesses a definite quantity of positive or negative electricity as an integral part of its constitution, and "every chemical action is fundamentally an electrical phenomenon," and "electricity is the first cause of all chemical activity." The varying degrees of chemical affinity were supposed to imply that the different substances were charged with *varying quantities* of electricity. When, say, sodium unites with oxygen to form the base sodium oxide, Na_2O ; and sulphur with oxygen to form the acid anhydride sulphur trioxide, SO_3 , primary compounds or compounds of the first order are formed. But the electrical attractions were not supposed to be always neutralized during the formation of these primary compounds, for the excess of positive or negative electricity caused a further attraction between the acidic and basic radicles, and resulted in the formation of compounds of higher orders.

The dualistic theory reigned supreme in chemistry for many years, but it was practically abandoned when Avogadro's hypothesis was generally

accepted, because Berzelius' theory was incompatible with such a comparatively simple reaction as $2\text{H}_2 + \text{O}_2 = 2\text{H}_2\text{O}$. According to Berzelius, the compound nature of oxygen is due to different electrical charges on the component atoms of the molecule. This does not agree with the supposed identity of the resulting two molecules of water, H_2O . Again, J. B. Dumas (1834) showed that the hydrogen atoms in compounds like CH_4 can be replaced one by one by atoms of chlorine. J. J. Berzelius had postulated that hydrogen is an electro-positive element, and chlorine electro-negative as exemplified in hydrogen chloride. Here, in Dumas' substitutions, a negative element can be exchanged for a positive element without fundamentally altering the chemical character of the resulting compounds. Facts like these brought Berzelius' theory into disfavour.

Faraday often expressed his conviction that "the forces termed chemical affinity and electricity are one and the same"; and that "chemical affinity is a consequence of the electrical attractions of particles of different kinds of matter." Contrary to Berzelius' assumption, Faraday proved that on electrolysis *definite and fixed quantities* of electricity are associated with the atoms of matter, although atoms of the same kind of matter in different compounds, on electrolysis, might be charged with different yet definite quantities of positive or of negative electricity. These statements are not antagonistic to "the kinetic theory of atoms" (p. 132).

§ 3. The Velocity of Electrolytic Conduction.

The conduction of electricity through electrolytes is utterly undistinguishable from metallic conductors, except for the action at the electrodes which is not part of true conduction at all.—J. T. SPRAGUE.

An electric current travels through an electrolytic solution as quickly as if the same current were sent through a copper wire of the same resistance, and the products of electrolysis appear simultaneously at both electrodes, however far apart the electrodes be placed. N. M. Hopkins (1905) passed a current through a tube 1500 cm. long, and through another tube 10 cm. long, and measured the time required for the current to pass by means of a chronograph sensitive to nearly 10,000 cm. per second. The tubes were filled with dilute sulphuric acid and fitted with electrodes—

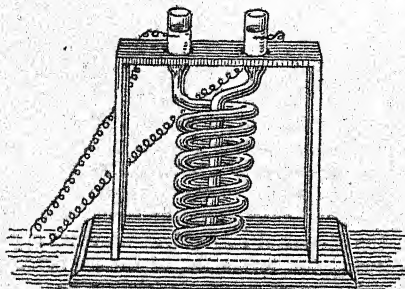


FIG. 134.—Velocity of Electrolytic Conduction.

the anode of copper, and the cathode of platinum. As soon as the current passes, bubbles of hydrogen appear at the cathode simultaneously with the blue colour of copper sulphate at the anode. The electrolyte 1500 cm. long conducted as quickly as an electrolyte 10 cm. long. The experiment can be illustrated by the apparatus sketched in Fig. 134, which almost explains itself. The long spiral tube contains the electrolyte

as in Hopkins' experiment. As soon as the circuit is closed, electrolysis begins. As A. E. Dolbear pointed out, "if the two terminals of an electric

circuit were on opposite sides of the Atlantic ocean, and a current were sent through the circuit, hydrogen would appear on one side and oxygen on the other, and in amounts defined by Faraday's law of electrolysis."

The known rates of diffusion of molecules in solutions are altogether too slow to allow the SO_4 which attacked the copper to have come from the same H_2SO_4 molecule as the hydrogen liberated at the cathode. Further, it is supposed that the electrical energy used in electrolysis is entirely expended in overcoming the resistance of the electrolyte, and no measurable quantity of work is needed for tearing apart the components of the decomposing molecule. Hence it follows that the molecules of an electrolyte in solution must be in a condition to conduct the electric current immediately the necessary electrical stress is applied to overcome the resistance of the liquid.

§ 4. The Effect of the Solvent.

Those bodies only are electrolytic which are composed of a conductor and of a non-conductor.—W. A. MILLER.

The more care taken in the purification of water, the less does it conduct electricity, and consequently, it is assumed that pure water is a non-conductor, in spite of the fact that perfectly non-conducting water has not yet been made. Pure dry liquid hydrogen chloride, like water, appears to be a non-conductor. A mixture of water and hydrogen chloride is an electrolyte. Hence it is inferred that the electrolytic conductivity of a solution is a property of solvent and solute, and not a property of either constituent alone. Solutions of dry hydrogen chloride in some solvents—*e.g.* dry benzene or chloroform—conduct electricity so feebly, if at all, that they are said to be non-conducting, and solutions of some substances in water conduct no better than water itself—*e.g.* solutions of sugar or alcohol in water. Hence also, it follows: the electrolytic conductivity of a solution depends upon some specific relation between the solvent and the solute. In a general way, aqueous solutions of acids, bases, and salts conduct electricity, and these substances are often called electrolytes, not because the salt conducts the current, but because their aqueous solutions conduct the current electrolytically. Some fused salts—*e.g.* silver chloride—conduct electrolytically, *e.g.* with fused silver chloride and silver electrodes, silver is dissolved at the anode and deposited at the cathode, so that the total amount of silver chloride remains constant; with carbon electrodes, silver is deposited on the cathode, and chlorine is evolved at the anode.

§ 5. The Ionic Hypothesis.

In framing hypotheses we must see that they agree with facts; in other respects they may be as inconceivable (not self-contradictory) as any fairy tale.—M. M. P. MUIR.

Let us learn to dream, then perhaps we shall find the truth.—A. KEKULÉ.

The main facts so far established by our discussion of the phenomena attending electrolysis may now be summarized:

- (1) Electrolytes in solution conduct electricity, and the process of electrical conduction is attended by a splitting of the molecules

of the solute into anions and cations; the anions appear at the anode, and the cations at the cathode. The separation of a certain number of anions at the anode is simultaneously attended by the separation of a chemically or electrically equivalent number of cations at the cathode. During electrolysis, the anions and cations appear to be discharged electrically, because electrically neutral molecules appear as secondary products of the electrolysis.

- (2) The anion which separates at the anode is not necessarily derived from the same molecule as the cation which appears at the cathode.
- (3) Solvent and solute together make a conducting medium, since as a rule neither solvent nor solute alone shows a marked capacity for conducting electricity.
- (4) No measurable time is needed to put an aqueous solution in a condition to conduct the current. Immediately the necessary difference of potential appears at the electrodes the process of electrolysis begins.
- (5) Osmotic pressure and related phenomena show that electrolytes in dilute solution have what seems to be a molecular weight, which suggests that the "ordinary" molecule of the electrolyte dissolved in certain solvents is dissociated into two parts.

As a trial hypothesis it may be assumed that the mere presence of the solvent leads to the fission of the molecules of the electrolyte into sub-molecules, each of which is charged with a definite amount of positive or negative electricity equivalent to 96,540 coulombs per chemical equivalent. The solution does not itself appear to be electrically charged, and hence it is assumed that equal quantities of positive and negative electricity are developed by the rupture of the molecules of the electrolyte during the process of solution. Solutions of electrolytes are supposed to normally contain a definite proportion of the sub-molecules charged with electricity. By a modification of Faraday's definitions (p. 38), the "sub-molecules" are called ions, and consequently: ions are atoms or groups of atoms which carry a positive or negative charge of electricity, and they are formed by the dissociation of the electrolyte in the solution. Each molecule, on dissociation, furnishes two kinds of ions with equal and opposite charges of electricity. Consonant with Faraday's laws, it is further assumed that each ion carries a definite charge of electricity (96540 coulombs)—a dyad atom carries two such charges; a triad atom, three charges, etc., but never a fraction of a charge. To avoid confusing the phenomenon of dissociation in which the products are not charged electrically with the dissociation of a molecule into electrically charged ions, the term ionization is reserved for the latter phenomenon. The ionization of hydrochloric acid is represented in symbols: $\text{HCl} \rightleftharpoons \text{H}^+ + \text{Cl}^-$; and of sodium chloride: $\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}^-$. This startling hypothesis appears so incredible and so opposed to the instinct, "common sense," or prejudices of the chemist that it has been assailed by much wholesome criticism—particularly by H. E. Armstrong. For instance, it is asked:

1. *In view of the great chemical activity of metallic sodium in contact with water, is it profitable to postulate the existence of the element sodium in contact with water without chemical action?* This objection is said to "rest

on a misunderstanding," because *electrically charged ions* of sodium in an aqueous solution of sodium chloride are very different from *neutral atoms* of metallic sodium. The ions of sodium carry large charges of electricity. It is urged that "chemists know practically nothing about the properties of atoms carrying large charges of electrical energy," and also that "the chemical activity of an 'atom' of sodium charged with its 96,540 coulombs of electricity is much less than a neutral atom of sodium." In other words, the presence of the electrical charge on the sodium ion keeps the ordinary chemical activities of the atom in abeyance. This means that whenever a chemical difficulty arises in the application of the ionic hypothesis the assumption is made that "neutral atoms or atomic groups and ions are different substances," because the properties of a substance are determined as much by the energy it contains as by the kind of matter. In this way, the ions have been invested with such imaginary properties as may be needed to keep the ionic hypothesis consistent with facts.

When the atoms or molecules of certain gases exist free in a liquid the gases can generally be removed, but when, say, sodium chloride is ionized, $\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}^-$, we are told that the chlorine does not escape because the chlorine ions, in virtue of their charge, cannot leave the liquid. Tolman, however, whirled aqueous solutions of sodium, potassium and hydrogen iodides in tubes in a powerful centrifugal machine, and found that the ends of the tubes acquired charges of opposite kind. The extreme ends of the tubes acquired a negative charge, attributed to the accumulation of the heavier positively charged iodine ions. The evidence is not complete, since electrical effects can be produced by rapid motions in air.

2. *Compounds like mercuric chloride, very prone to thermal dissociation, are not readily ionized; while compounds like calcium chloride which resist thermal dissociation are readily ionized. Would not the ionic hypothesis predict the converse phenomena?* Mercuric chloride is very volatile and readily dissociates into its elements by heat; calcium chloride, on the contrary, does not readily volatilize or dissociate except at very high temperatures, yet it is said that the latter is readily ionized in solution while the former remains all but unchanged. Here again it is said that totally different phenomena are confused, and that the forces which produce ionization are quite different from those which produce thermal dissociation.

3. *Bodies carrying electrical charges of opposite sign are attracted and cling to one another; if therefore a mobile solution contains "free and independent" ions carrying enormous electrical charges of opposite sign, how can the charged ions remain more than momentarily free?* It is assumed that a certain proportion of the molecules of the solute are continually breaking down into free (charged) ions, and a certain proportion of the ions are continually recombining to form ordinary molecules, the result is that the ratio between the number of free ions and paired ions (molecules) remains unchanged. This statement, of course, does not answer the perplexing question. Attempts have been made to refer the difficulty to the specific insulating properties—the so-called dielectric constant—of the solvent. The action of the solvent has been compared with the function of the glass in a charged Leyden jar. This agrees with the non-conducting qualities of pure water, but experiments have shown that the relation between the insulating properties of a solvent and its ionizing properties

is not an adequate and sufficient explanation of the observed facts. The two phenomena do not always vary concomitantly. A satisfactory answer to the question, therefore, has not yet been found.

4. *If an ionized salt, say, sodium chloride, is present in solution as a mixture of Na^+ and Cl^- ions, it might be thought possible to separate the two components by diffusion or by some other mechanical process.* S. Arrhenius answers that the great electrostatic attraction of the oppositely charged ions prevents any marked diffusion. W. Nernst, however, has shown that the concentration currents produced when, say, a solution of sodium chloride is carefully covered with a layer of water, leads to the conclusion that the greater mobility of the chlorine ions charges the upper layer negatively, and the lower layer positively, so that a current of electricity can be obtained by placing the two layers in electrical contact.—See “Concentration cells.”

5. *Salts which form solid compounds with different proportions of water of crystallization have different solubilities; is it not then more reasonable to suppose that the molecules of the solvent exist in solution as definite hydrates?* The ionic hypothesis answers that only a definite proportion of the salt is ionized, and this part is proportionally less as the concentration of the solution increases. As a rule, in a saturated solution only a small fractional part of the salt is supposed to be ionized.

6. *When a compound is formed from its elements with the loss of energy, the compound cannot be resolved into its elements unless energy be supplied. It is therefore pertinent to inquire: What is the source of the energy which leads to the fission of the molecule into ions carrying equal but opposite charges of electricity?* Here, again, it is necessary to reiterate that the ionic hypothesis refers not to the separation of a compound into its original constituents, but into charged ions; and it is interesting to observe that molecules of sodium chloride, etc., which appear to be very stable when dry, react with great facility when in solution. A little heat is supposed to be evolved during the ionization of many (not all) electrolytes, and the process of ionization is then presumably accompanied by an exothermal reaction which more than compensates for the energy needed for the fission of the molecule into oppositely charged ions. J. D. van der Waals (1891) expressed the idea that ionization is primarily due to the affinity of ions for the solvent, and that the heat of hydration of the ions furnishes the energy needed for the ionization of the salt. To this, D. Konowalow (1893) adds, “only those solvents which react chemically with the solute furnish solutions which conduct electricity.” G. T. Beilby considers that the ionization is essentially a mechanical operation, the result of the kinetic activity of the solute molecules, for in a dilute aqueous solution of, say, hydrogen chloride, each molecule of the solute is surrounded by and at the mercy of some millions of water molecules all in a state of intense activity, and the rude mechanical jostling to which the molecule of hydrogen chloride is subjected will naturally tend to break it up into simpler parts mechanically more stable.

These answers, based on circumstantial evidence, are not altogether satisfactory, but they have made it almost certain that if ions are formed at all, the ions do not usually exist alone in the solution, but that each ion forms a complex by association with a definite number of molecules of the solvent; or, as R. Abegg expressed it, the degree of ionization

depends on the capacity of the ions to unite with the molecules of the solvent. The ionic theory primarily assumes that the apparent *number* of solute "molecules" is increased by ionization, and it makes very little difference to the applications of the ionic theory whether it be assumed that each ion is isolated as a distinct individual, or whether each ion forms a complex with the molecules of the solvent. The number of ions is the same in both cases. The explanation of the phenomena, particularly when solutions other than water are considered, is beset with many difficulties on account of the unknown relations between the solvent and solute. So much is this the case that M. le Blanc (1907) considered "it to be very fortunate for the advance of electro-chemistry that such complications are generally, though not always, absent in the case of aqueous solutions. It is due to this fact that it has been possible to deduce simple laws from the study of aqueous solutions." On the other hand, O. N. Witt said: "Water is a very complicated substance, and the process of solution in that liquid must be attended by very great exceptions from the simple rules which exist for solvents not so complicated." It has therefore been a great mistake to study aqueous solutions and then other solutions, and the chapters in treatises on physical chemistry entitled "theory of solution," would be rightly labelled "theory of aqueous solution."

Several attempts have been made to work out a consistent explanation of the fundamental facts without a theory of charged ions, but with hypotheses based upon the formation of imaginary molecular complexes by a reaction between polymerized solvent and the molecules of the solute. The difficulties, however, are discussed in special text-books.

These controversial matters are mentioned in this elementary book to emphasize the fact that an explanation of a phenomenon may contain part of the truth, and yet not "the whole truth, and nothing but the truth." In that case, we try the hypothesis by the test indicated in the first chapter, and ask: Is the hypothesis useful? The answer is that the ionic hypothesis has done good work, and it promises to do more. An hypothesis is not always to be discarded as a first approximation because troublesome exceptions crop up from time to time. Newton's theory of gravitation, for instance, appeared to be afflicted with such blemishes—particularly in its early days; so was the theory of opposing reactions once considered to be unreasonable folly; and the present-day theory of light seems highly absurd when it is remembered that it is based on the existence of an ether which is of the highest elasticity and denser than steel. In spite of important difficulties, we shall now try how the ionic hypothesis fits in with a few important phenomena.

History.—R. Clausius (1857) first suggested that the molecules of a solute are ionized when dissolved in the solvent, but Clausius appears to have had some "chemical" misgivings, for he added that the ionization only affected an infinitesimally small fraction of the total number of dissolved molecules. As the ions are discharged at the electrodes during electrolysis, more molecules are ionized. The un-ionized molecules keep the electrolyte constantly supplied with a definite number of ions. The ions conduct the current; the "undissociated" molecules are inactive. Further, at any given temperature, there is a constant relation between the number of un-ionized molecules and the number of ions. W. Giese (1882), and S. Arrhenius (1884), more bold, or less cautious, than Clausius,

centimetre apart, form the electrodes. Otherwise expressed, the equivalent conductivity represents the conducting power of a layer of the solution 1 cm. thick, and containing one gram-equivalent of the substance in solution. If the conductivity be referred to a gram-molecule, and not a gram-equivalent, it is termed the molecular conductivity of the solution, symbolized μ .

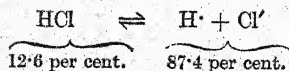
It is now assumed that the number of ions which take part in conducting the electric current at any particular concentration of the solution is proportional to the equivalent conductivity, λ , of the solution. If the number of ions in a given solution be doubled, the conductivity will also be doubled. Consequently, if x represents the fraction of a gram-molecule which is dissociated into ions when the solution occupies v litres, we have, at dilution v , the conductivity— $\lambda_v = kx$, where k is the constant of proportion. At infinite dilution, the whole gram-molecule is supposed to be ionized, and consequently, $x = 1$, and therefore the conductivity, λ_x , at infinite dilution, is $\lambda_x = k$. Substitute this value of k in the preceding equation, and we get

$$x = \frac{\lambda_v}{\lambda_x}$$

which, by hypothesis, means that the fractional number of molecules ionized in a solution is numerically equal to the equivalent conductivity of the solution divided by the equivalent conductivity of the solution at infinite dilution; or,

$$\text{Degree of ionization} = \frac{\text{Number of molecules ionized}}{\text{Total number of molecules}} = \frac{\lambda_v}{\lambda_x}$$

The above formula enables the electrical conductivity of a salt to be expressed in terms of the degree of ionization of the salt in solution. Thus, the equivalent conductivity of a solution of hydrochloric acid is 305.4, and the same acid at infinite dilution has the equivalent conductivity 349.3. Hence the degree of ionization is $305.4 \div 349.3 = 0.874$ per gram molecule, or 87.4 per cent. ionization. This means that 12.6 per cent. of the molecules in the solution are present as un-ionized neutral molecules, HCl; and 87.4 per cent. of the molecules are present in the ionic form $\text{H}^+ + \text{Cl}^-$. Or,



The *percentage ionization* must not be confused with the *absolute concentration of the ions*. The former may be the greater in dilute solutions, and the latter greater in concentrated solutions. The ionic hypothesis thus assumes that an aqueous solution of hydrochloric acid contains three distinct kinds of "solute molecules," electrically charged molecules (hydrogen and chlorine ions), and neutral hydrogen chloride molecules.

Table XVII., on next page, shows the degree of ionization of normal solutions (unless otherwise stated) of a few typical acids, bases, and salts selected merely for illustrative purposes.

The effect of increasing the concentration of a solution is to increase the internal friction. This retards the movements of the ions and thus diminishes the conductivity more rapidly than would occur if the results were not affected by this disturbing factor. As the concentration decreases

TABLE XVII.—DEGREE OF IONIZATION OF SOME TYPICAL ACIDS, BASES, AND SALTS.

ACIDS.		BASES.		SALTS.	
Acid.	Fraction ionized.	Base.	Fraction ionized.	Salt.	Fraction ionized.
Nitric acid (62 %) .	0.096	Potassium hydroxide	0.77	Potassium chloride	0.74
Nitric acid (dil.) .	0.820	Sodium hydroxide .	0.73	Ammonium chloride	0.75
Sulphuric acid (dil.)	0.510	Lithium hydroxide .	0.63	Potassium nitrate .	0.64
Carbonic acid ($\frac{1}{10}$ N)	0.002	Ammonium hydroxide	0.01	Zinc sulphate . . .	0.24
Hydrosulphuric acid ($\frac{1}{10}$ N) . .	0.001	Calcium hydroxide ($\frac{1}{10}$ N)	0.90	Copper sulphate . .	0.22
Perchloric acid ($\frac{1}{10}$ N)	0.880	Barium hydroxide ($\frac{1}{10}$ N)	0.92	Silver nitrate . . .	0.58
Acetic acid (0.1 N)	0.013			Barium chloride . .	0.57
				Potassium sulphate	0.24

the friction diminishes; and, with the more dilute solutions, the effects of internal friction can be neglected. The two factors—internal friction and conductivity—do not change with dilution in the same way, and, in consequence, the conductivity may increase with increasing dilution; reach a maximum; and then decrease with increasing dilution. This is the case, for instance, with sulphuric acid, where the maximum conductivity occurs when 30 per cent. of acid is present. The application of the ion theory to *concentrated* solutions is beset with many difficulties, and consequently the theory has been mainly developed from results obtained with *dilute* solutions. Similar remarks apply to solvents other than water.

Illustration of the different conductivities of electrolytes.—The difference in the conductivity of, say, hydrochloric, sulphuric, chloracetic, and acetic acids, containing $\frac{1}{10}$ gram-equivalent of the respective acids per litre, is well shown by W. R. Whitney's apparatus (1900), illustrated in Fig. 136. Electrodes are fixed in four vertical tubes. The upper electrodes are connected with one terminal (as shown by the dotted line), which is connected with the lighting circuit. Each of the lower electrodes is connected with one terminal of an ordinary incandescent lamp. The other terminals of the lamps are connected with another terminal which is connected with the lighting circuit (say an alternating current, 110-volt). The upper electrodes can be moved up or down; the lower electrodes are fixed. The tubes are filled

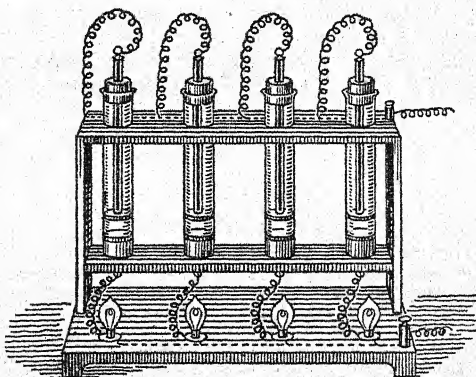


FIG. 136.—Different Conductivities of some Acids.

with the four acids just indicated, and the electrodes are adjusted the same distance apart. The current is switched on. The lamp connected with the hydrochloric acid tube glows brightest; that with the sulphuric acid comes next; that with chloracetic acid next, and the one connected with the acetic acid tube scarcely glows at all. The electrodes can now be adjusted until the lamps all glow equally bright. It will be found that the distances between the electrodes are approximately as 100 : 85 : 15 : 1 when the acids are arranged in the order named above. These numbers give an approximate measure of the conductivity of these acids, and also, by hypothesis, of their degrees of ionization.

Strong and weak acids and bases.—The terms "strong" and "weak" are sometimes applied to the acids and bases, and these terms refer to the conductivity or the degree of ionization in aqueous solution of moderate dilution. There is no real line of demarcation between the two. Acids like carbonic and hydrosulphuric acids, and bases like ammonia, are weak. Their degree of ionization is less than one per cent. If the degree of ionization exceeds 70 per cent. the acid is undoubtedly strong.¹ Most of the salts are highly ionized, even at moderate dilutions, but there are many exceptions, *e.g.* mercuric chloride; the cadmium halides and mercuric cyanide are but slightly ionized in moderately dilute solutions.

§ 7. The Migration of Ions.

It is impossible to get one kind of ion liberated at one electrode without having a precisely equivalent quantity of an oppositely charged ion appearing at the other electrode; it is impossible to have a procession of positive atoms through a liquid without a corresponding procession of negative ones. In other words, an electric current in a liquid necessarily consists of a flow of positive electricity in one direction combined with a flow of negative electricity in the opposite direction.—O. J. LODGE.

Although the quantities of anions and cations liberated at the electrodes during electrolysis are always strictly equivalent, nevertheless, the rates at which the concentration of the electrolyte changes about anode and cathode are not the same. This is illustrated by the following experiments:

Mix a hot 5–10 per cent. aqueous solution of gelatine with a little cupric chloride solution, and pour the blue solution, while hot, into a warm U-tube with long legs until the U-tube is a little over half full. Take care to introduce as few air bubbles as possible. Let the mixture solidify by cooling. When cold, slip a rubber ring over each leg of the tube so as to mark the level of the gelatine in each leg. Pour a dilute solution of sodium chloride into each leg; dip a piece of platinum wire into the solution of sodium chloride; and pass a current of electricity through the contents of the U-tube while the latter is immersed in a freezing mixture,² Fig. 137, to prevent the softening of the gelatine during the passage of the current. The gelatine simply serves to prevent mechanical movements of the liquid during the passage of the current. The presence of the gelatine does not otherwise interfere with the electrolysis. The experiment is more interesting if a second U-tube be prepared with potassium dichromate in place of copper chloride, and connected in series with the copper chloride tube as indicated in Fig. 137. A current of electricity is then passed through the tubes *via* the platinum wire electrodes, each of which dips in a dilute solution of sodium chloride above the gelatine. After the current has passed some time, the blue

¹ Electrolytes like solutions of sodium chloride are good conductors, and some solutions with a conductivity midway between good conductors and non-conductors are sometimes called semi- or half-electrolytes.

² Say, ammonium nitrate and cold water.

solution rises past the level of the solid gelatine towards the cathode, and descends below the level of the gelatine on the anode side. The gelatine has not moved, and it appears as if the current "drives" the copper chloride molecules towards the cathode; the movements in the other tube are in the converse direction, so that it appears as if the current "pulls" the molecules of the potassium dichromate towards the anode. One tube with copper dichromate can be used in place of copper chloride and potassium dichromate. The gelatine is green, but a blue coloration rises towards the cathode, and a yellow coloration rises towards the anode.

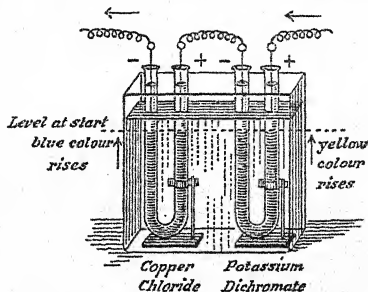


FIG. 137.—Migration of Ions to the Electrodes.

The ionic theory interprets these experiments of A. A. Noyes and A. A. Blanchard (1900), by assuming that in the case of cupric chloride, blue copper ions travel towards the cathode and colourless chlorine ions towards the anode; in the second experiment, that yellowish Cr_2O_7 -ions travel towards the anode and colourless potassium ions towards the cathode; and in the third experiment, blue copper ions travel towards the cathode and yellow dichromate ions towards the anode. The fact observed is that the electrolysis of the coloured solutions occurs at the boundary surfaces between the gelatine and the supernatant solution. M. Faraday described several experiments illustrating the phenomenon: the surfaces of separation of liquids in contact act as electrodes to each other, and separation may there occur, just as at a plate.

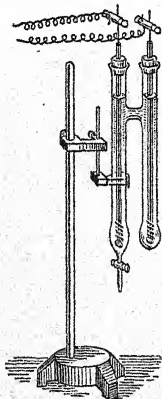


FIG. 138.—Determination of the Speed of Ionic Migration.

By a modification of these experiments, it is possible to measure the rates at which the concentration of the solution changes about the electrodes, or, in the language of the ionic theory, the rates at which the anions of copper, etc., drift towards the electrodes. For example, if a solution of silver nitrate of known concentration be electrolyzed between silver electrodes in an apparatus similar to that illustrated in Fig. 138, the only change in the solution is a transfer of silver from the anode to the cathode, and a change in the concentration of the silver salt round the two electrodes; for the apparatus is constructed so as to reduce the mechanical convection of the dissolved salt to a minimum. The change in the concentration of the solution, after a few hours' electrolysis, can be measured by withdrawing about half the solution from the apparatus, *via* the stopcock, and determining the amount of silver in the solution by analysis. From the results, numbers can be obtained which are supposed to represent the speeds of migration of the anions and the cations. The following experiment, due to W. Hittorf (1853), serves to illustrate the principle.

Hittorf's experiment.—A solution of silver nitrate containing one part

of silver to 49.44 parts of water was electrolyzed for nearly an hour in a cell with silver electrodes. Silver dissolved from the anode and a similar quantity deposited on the cathode. The concentration of the whole solution remained unchanged, but the concentration of the solution about the cathode decreased while that about the anode increased. In the cathode compartment, Hittorf found

Silver before electrolysis	0.7162 gram
Silver after electrolysis	<u>0.5862</u> "
Loss	0.1300 "

The solution about the cathode thus lost 0.1300 gram of silver, and the solution about the anode must have increased by this amount owing to the action of an equivalent quantity of nitric acid on the silver electrode. At the same time, by the simultaneous interposition of a silver voltmeter in the circuit, it was found that sufficient electricity had passed through the electrolyte to deposit 0.2470 gram of silver at the cathode. If no silver ions had passed from the anode chamber, the quantity of silver in the anode chamber would have increased by 0.2470 gram owing to the migration of NO_3' ions from the cathode chamber. The observed increase was only 0.1300 gram of silver; hence 0.2470 less 0.1300 gram; in all, 0.1170 gram of Ag^+ ions migrated from the anode chamber to the cathode chamber, while the cathode chamber simultaneously lost 0.1300 gram of silver due to the deposition of 0.2470 gram of silver on the cathode. Hence, since the relative speeds of the ions are proportional to the fall of the concentration about the oppositely named electrodes,

$$\frac{\text{Loss in cathode chamber}}{\text{Loss in anode chamber}} = \frac{0.1300}{0.1170} = \frac{\text{Speed of anion, } \text{NO}_3'}{\text{Speed of cation, } \text{Ag}^+}$$

Or the rate of migration of the anions is to the rate of migration of the cations as 130 : 117; or the NO_3' ions migrate 1.1 times as fast as the Ag^+ ions.

According to F. Kohlrausch's second law, the molecular conductivity at infinite dilution is the sum of two factors, one dependent on the anion and the other on the cation, so that $\mu_\infty = v + v'$, where v denotes the speed of migration of the cations, and v' that of the anion. Again, the proportion of the current carried by the anion is $v'/(v + v')$, and the proportion carried by the cation $v/(v + v')$, where $v + v'$ represents the rate at which the ions separate, or the joint velocity of the two ions. The relative proportion of the current carried by each ion, is v/v' ; this number or ratio is called the transport number.

By measuring the rate of rise of the blue colour in the experiment, Fig. 137, the absolute velocity of copper ions can be determined under standard conditions, and in that way, with other solutions, a series of numbers have been obtained which are called the speeds of migration of the different ions. The speed of migration so defined is a specific property of each ion, and is independent of the other ions present. The hydrogen ion travels faster than any other. The speeds are increased by using currents of greater electromotive force. Hence the rate of motion of any given ion is determined (i) by the intensity of the electrical pressure, which directs or drives the ions to the electrodes; (ii) by the damping effect of the liquid on the motion of the moving ions. Under constant

conditions, at 18°, with a difference of potential of one volt between the electrodes, the speeds of migration of some ions are :

Anions	OH'	Cl'	I'	NO ₃ '	ClO ₃ '	
Speeds	5.6	2.12	2.19	1.91	1.70	cm. per hour.
Cations	Cs.	Rb.	K.	Na.	Li.	
Speeds	2.32	2.32	2.05	1.26	1.11	cm. per hour.

The heaviest ions, that is, the ions with the greatest "atomic" weights, here appear to move fastest. This has been supposed to be due to the slower-moving ions dragging along with them a number of molecules of the solvent. As a matter of fact, if the speeds of migration of the different ions be plotted on squared paper, against the atomic weights, a periodic curve is obtained.

In these experiments the *fact* observed is the changing molecular concentration of the solution about the anode and cathode during electrolysis; the extraordinary *hypothesis* is that during the passage of the current the anions and cations move in the same electrolyte with different velocities, and yet the anions and cations are given off at the respective electrodes at the same time !

§ 8. "Abnormal" Osmotic Pressures and Ionization.

It is natural to assume that substances which in aqueous solution give osmotic pressures which are too great are dissociated.—S. ARRHENIUS.

We are now in a position to resume our study of the abnormal osmotic pressures furnished by solutions of electrolytes. Just as the abnormally high vapour densities exhibited by, say, hydrogen fluoride were traced to the polymerization of the molecule, $2\text{HF} \rightleftharpoons \text{H}_2\text{F}_2$, and the abnormally low vapour densities of iodine above 700° were traced to the dissociation of the iodine molecule, $\text{I}_2 \rightleftharpoons 2\text{I}$, so Arrhenius argued that the molecules of salts giving an abnormally high osmotic pressure must be dissociated. Suppose that one molecule of an electrolyte furnishes m ions, and further let x denote the fraction ionized when a gram-molecule of the electrolyte is dissolved in water. The solution will then contain $(1-x)$ non-ionized molecules, and mx ions. The total number of individual molecules in the solution—that is, electrically charged molecules (ions) and neutral molecules—will be $(1-x) + mx$. As in our previous study of solutions, let n denote the total number of individual molecules in a given solution. Then $n = 1 + mx - x$. The numerical value of n , as we have seen, can be determined from conductivity data, and from osmotic pressure and related phenomena—freezing and boiling point determinations. If the value of n so determined be divided by the value of n calculated on the assumption that no ionization occurs, the value of x and accordingly also the value of n can be computed. For example, the solution of hydrochloric acid just studied gives $x = 0.874$, and $m = 2$. Hence $n = 1 + (m-1)x$ becomes $n = 1.874$. Hence every 100 molecules of HCl furnish the equivalent of 187.4 individual molecules. If the electrolyte had been non-ionized, n would have been unity; and if completely ionized, n would have been 2.

EXAMPLES.—(1) The boiling point of a solution of 3.400 grms. of barium chloride, BaCl_2 , in 100 grms. of water is 100.208°. What is the ionization of the

solute? From the above, $1 + mx - x$, it follows that there will be $1 + 3x - x$, or $1 + 2x$ molecules in the solution, *i.e.* Ba^+ and 2Cl^- . If there was no ionization, the molecular weight of the barium chloride would be 208; the observed value is 85. Hence one molecule of barium chloride gives the equivalent of $208/85 = 2.447$ molecules. Hence $2.447 = 1 + 2x$, or $x = 0.723$, or 72.3 per cent. of the salt is ionized.

(2) A solution of 11.07 grms. of barium nitrate in 100 grms. of solvent raised the boiling point 0.466° , show that 55.8 per cent. of the salt is ionized.

(3) A solution of 0.3665 grm. of sodium chloride in 100 grms. of water freezes at -0.221° , show that 89.2 per cent. of the salt is ionized.

A comparison of the values of n calculated from osmotic pressure, freezing point, and electrical conductivity data is indicated in Table XVIII.

The numbers in the last three columns show that the values determined by independent processes are strikingly concordant; and it is therefore inferred that the abnormal osmotic pressures indicated in Table IX., p. 263, arise from the more or less complete ionization of the electrolytes in aqueous solution.

Modes of ionization.—The ionization of some of the multivalent electrolytes— H_2CO_3 ; H_2SO_4 ; BaCl_2 ; CdCl_2 ; K_2SO_4 ; CuSO_4 ; etc.—in moderately dilute solutions appears to furnish complex ions. Thus, cadmium chloride, CdCl_2 , not only furnishes $\text{Cd}^{++} + 2\text{Cl}^-$, but also $\text{Cd}^{++} + \text{CdCl}_4^{--}$; sulphuric acid, H_2SO_4 , not only gives $2\text{H}^+ + \text{SO}_4^{--}$, but also $\text{H}^+ + \text{HSO}_4^-$; carbonic acid, H_2CO_3 , gives $2\text{H}^+ + \text{CO}_3^{--}$, and $\text{H}^+ + \text{HCO}_3^-$; copper sulphate, CuSO_4 , not only gives $\text{Cu}^{++} + \text{SO}_4^{--}$, but also $\text{Cu}_2\text{SO}_4^{++}$ and $\text{Cu}(\text{SO}_4)_2^{--}$; etc. If, however, the solutions be still further diluted, the complex ions break down into simpler ones. Hence the ionization of concentrated polybasic acid like H_2SO_4 proceeds in stages first $\text{H}_2\text{SO}_4 = \text{H}^+ + \text{HSO}_4^-$; followed by $\text{H}^+ + \text{HSO}_4^- = 2\text{H}^+ + \text{SO}_4^{--}$.

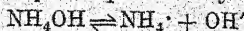
TABLE XVIII.—MOLECULAR WEIGHTS OF SOME ELECTROLYTES IN SOLUTION.

Salts.	Molecular concentration.	Values of n .		
		Osmotic pressure.	Freezing point.	Conductivity.
Calcium nitrate, $\text{Ca}(\text{NO}_3)_2$.	0.18	2.48	2.47	2.46
Magnesium sulphate, MgSO_4 .	0.38	1.25	1.20	1.35
Strontium chloride, SrCl_2 .	0.18	2.69	2.52	2.51
Potassium chloride, KCl .	0.14	1.81	1.86	1.86
Lithium chloride, LiCl .	0.13	1.92	1.94	1.84
Magnesium chloride, MgCl_2 .	0.19	2.79	2.68	2.48

§ 9. Equilibrium between Ionized and Non-ionized Solute.

The evidence is so unambiguous and convincing that ions and some molecules combine with more or less of the solvent, that it seems that it can now be accepted as a fact of science.—H. C. JONES.

Reference has previously been made to the assumption that the molecules of an electrolyte, when dissolved in water, are ionized; that the ions, at the same time, recombine to form neutral molecules; and that equilibrium will ensue when the speeds of the two opposing reactions—ionization and de-ionization—are equal. Consider the ionization of ammonium hydroxide, NH_4OH , represented by



Here the process of ionization bears some analogy with the dissociation of iodine by heat: $I_2 = I + I$. Let $[NH_4OH]$ denote the concentration of the ammonium hydroxide; $[NH_4']$ the concentration of the ammonium ion; and $[OH']$ the concentration of the hydroxide ion. Then, applying the principle of opposing reactions, the condition for equilibrium is:

$$\frac{[NH_4'] \times [OH']}{[NH_4OH]} = K$$

If this theory applies to ions, the numerical value of the equilibrium constant, now called the ionization constant, remains unchanged whatever be the concentration of the solution. This principle is sometimes called W. Ostwald's dilution law.

EXAMPLE.—In a solution containing 0.125 gram-molecule of NH_4OH per litre, the equivalent conductivity shows that 0.0135 gram-molecule is ionized, and hence, $0.0135 \times 0.125 = 0.0017$ represents the molecular concentration of the ammonium hydroxide which is ionized. This number thus represents the concentration of the NH_4' ions. But every NH_4' ion is accompanied by one OH' ion, and accordingly 0.0017 also represents the concentration of both the NH_4' and the OH' ions. Hence, from Ostwald's dilution law, $0.0017 \times 0.0017 \div (0.125 - 0.0017) = 0.000023$. This last number represents the value of the ionization constant for an $\frac{1}{8}$ -normal solution of ammonium hydroxide.

If the solution be diluted, the *percentage amount* of ionization increases (although, of course, the *actual* concentration of the ions must decrease); and with more concentrated solutions, the percentage amount of ionization decreases. Table XIX. represents values for the ionization constant for solutions of ammonium hydroxide of different strength.

TABLE XIX.—EFFECT OF DILUTION ON THE IONIZATION OF AQUEOUS AMMONIA.

Ammonium hydroxide. (Gram-molecules per litre.)	Proportion ionized.	Molecular concentration of NH_4' and of OH' ions. (Gram "ions" per litre.)	Molecular concentration of non-ionize NH_4OH . (Gram-molecules per litre.)	K
1.0000	0.0047	0.0047	1.0000-0.0047	0.000023
0.1250	0.0135	0.0017	0.1250-0.0017	0.000023
0.0159	0.0376	0.0006	0.0159-0.0006	0.000023
0.0039	0.0754	0.0003	0.0039-0.0003	0.000023

The constancy in the value of K means that although the last-named solution of ammonium hydroxide is nearly 300 times more dilute than that named first, and the degree of ionization of the last is nearly 16 times as great as the first, the expression represented by K , deduced on the supposition that the process of ionization follows the rule for opposing reactions, is constant within the limits of experimental error.

When this method is applied to strong acids and bases—that is, acids and bases more highly ionized than ammonium hydroxide—the results are not nearly so good, and the most satisfactory explanation of the discrepancy turns on the assumption that the dissolved substance unites with the solvent so that in the more concentrated solutions part of the liquid in which the substance is dissolved no longer functions as a solvent because it

is in combination with ions and non-ionized molecules of the solute. This assumption is at the base of the solvate theory of solutions advocated by H. C. Jones.

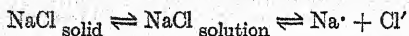
§ 10. The Solubility Law.

In the case of an aqueous solution of sodium chloride, containing, say, 58.5 grams, that is, one gram-molecule per litre, we have: $\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}'$, where 68 per cent. of the salt is ionized. The condition of equilibrium, according to the dilution law, is

$$\frac{[\text{Na}^+][\text{Cl}']}{[\text{NaCl}]} = K; \quad \frac{0.68 \times 0.68}{0.32} = 1.44 = K$$

If either Na^+ or Cl' ions be added ¹ to the solution—say, by adding hydrogen chloride, making the concentration of the Cl' ions 0.75 instead of 0.68—then, in order to preserve the constancy of the ratio 1.44, the concentration of the Na^+ ions must be diminished. This can only occur by the union of some Na^+ and Cl' ions to form NaCl until the ratio K is again 1.44.

Solubility of mixtures with a common ion.—Sodium chloride is in equilibrium with its aqueous solution when, at a given temperature, the concentration of the substance in solution has a certain definite and constant value—the solubility of the substance at the given temperature. Since the sodium chloride in solution is partly ionized, there are two equilibria to consider; first, the relation between the non-ionized and the ionized salt in solution: $\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}'$ just indicated; and second, the relation between the non-ionized salt and the solid. If the solution be saturated, we have:



If the concentration of the Na^+ or the Cl' ions be augmented, some of the Na^+ and Cl' ions will recombine to form non-ionized NaCl as indicated above. Consequently, some sodium chloride will be precipitated or the solution will be supersaturated. Hence the solubility of a salt is usually diminished in the presence of another compound with a common ion. If the solution of the hydrochloric acid had been isohydric with the salt solution—i.e. if the number of chlorine ions per cubic centimetre had been the same—no alteration in the concentration of the ions would occur, and therefore no salt would be precipitated on mixing the solutions, provided no disturbing secondary action occurs.

Solubility products.—It is sometimes convenient to discriminate between the total or apparent solubility of a salt, and the amount of the non-ionized salt present in the solution. The latter is sometimes called the real solubility of the salt. In a saturated solution the real solubility, like the apparent solubility, must be constant. Hence, in the dilution law for sodium chloride: $[\text{Na}^+][\text{Cl}'] \rightleftharpoons K [\text{NaCl}]$, indicated above, the concentration $[\text{NaCl}]$ is invariable, K is constant, and consequently also the product of the two is constant. Therefore, we can write for saturated solutions:

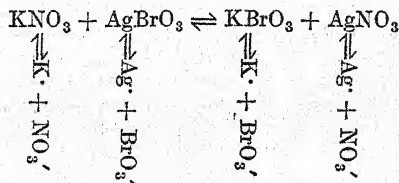
$$[\text{Na}^+][\text{Cl}'] = \text{Constant}$$

¹ If Cl' ions be added, they must necessarily be accompanied by an equivalent number of oppositely charged ions of, say, K' , from, say, a solution of potassium chloride; or of Na^+ ions from, say, sodium hydroxide: $\text{NaOH} \rightleftharpoons \text{Na}^+ + \text{OH}'$.

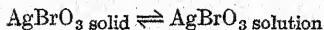
This relation means that in a saturated solution the product of the "molecular" concentrations of the ions is constant. This product is sometimes called the solubility product because, from what has been already stated, the product of the two ion concentrations determine the magnitude of the "real" solubility of the substance.

Familiar examples of this phenomenon are the precipitation of sodium or potassium chlorides from saturated solutions by passing hydrogen chloride through the solutions, or by adding a concentrated aqueous solution of the acid. The phenomenon is quite general. Barium chloride may be substituted for sodium chloride; again, nitric acid will precipitate barium nitrate from concentrated aqueous solutions; a nearly saturated solution of silver bromate will give a precipitate of silver bromate, if either silver nitrate or sodium bromate be added to the solution; sodium chloride added to a saturated solution of potassium chloride will lead to the precipitation of the last-named salt. There are a number of complications in special cases, thus: when a nearly saturated solution of sodium chloride is treated with alcohol, or with hydrogen chloride, the solvent combines with the added material, and less is available for the solution of the salt in question; the solute may form polymerized molecules in the solution, etc.

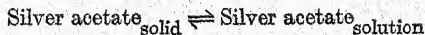
Solubility of mixtures with no common ion.—If potassium nitrate be added to a saturated solution of silver bromate, a number of molecules of silver nitrate and potassium bromate will be formed by double decomposition, and the solution will be in equilibrium when these four salts have attained a definite concentration, and each salt is itself ionized and is in equilibrium with the corresponding ions. The condition of equilibrium is therefore quite complex. It may be symbolized:



The net result is that the number of ionized and non-ionized molecules of silver bromate in the solution is lessened, and the equilibrium:



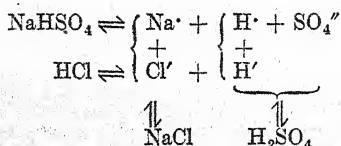
is disturbed. The original relation is restored by the passage of more silver bromate into solution. Similarly, when nitric acid is added to a saturated solution of silver acetate, some silver nitrate is formed, and the equilibrium:



can only regain its former value by the passage of more silver acetate into solution. Consequently, the solubility of a salt is often increased in the presence of a compound containing no common ion. As before, complications arise owing to the dehydration of the solution by the added substance, etc. The dehydrating action is illustrated in the familiar method of preparing hydrogen chloride by dropping concentrated sulphuric

acid into a concentrated hydrochloric acid. The sulphuric acid abstracts water, and thus diminishes the effective solvent, the hydrogen chloride is then evolved as a gas. The action is in part attributed to the repression of the ions of hydrogen chloride at the same time, and the consequent elimination of molecules from the water already saturated with molecules of the same kind. Similarly, an excess of sulphuric acid precipitates the di- or mono-hydrate of cadmium chloride almost completely from an aqueous solution—which hydrate depends on the amount of the sulphuric acid used. Cadmium, mercuric, and stannous chlorides and bromides behave similarly. The behaviour of many salts is thus not completely described by the “solubility product” law.

Rule for precipitation.—The solubility product of sodium chloride in solution is not very great, and, in consequence, if concentrated hydrochloric acid be added to a concentrated solution of sodium hydrogen sulphate, NaHSO_4 , the solubility product of sodium chloride may be exceeded, and that salt will be precipitated. The condition of equilibrium of the mixed solutions is :



Hence, if the product of the “molecular” concentration of any pair of ions (with equal and opposite electrical charges) in a solution be greater than the solubility product for the saturated solution formed by the union of these ions, that substance will be precipitated; and conversely, if a substance be present in excess, it will be dissolved if the product of the “molecular” concentration of any pair of ions (with equal and opposite electrical charges) in a solution be *less* than the solubility product for the saturated solution formed by the union of these ions.

§ 11. Acids and Bases according to the Ionic Hypothesis.

The dissociation theory enables us to place the whole subject of acidity upon a rational basis. Without this theory the subject would still be only so many empirically established, disconnected, and meaningless facts.—H. C. JONES.

Acids.—It will be remembered that Gerhardt defined acids to be “salts of hydrogen,” the ionic hypothesis expresses a similar idea another way: “all acids, when dissolved in water, furnish hydrogen ions.” Although many substances not usually called acids when completely ionized furnish hydrogen ions—*e.g.* potassium hydrogen sulphate, KHSO_4 , etc.—yet their acidic properties are due to the presence of H^+ ions, and consequently it has been said that “there is only one acid, and that is the H^+ ion,” and that “the two terms ‘acidity’ and ‘hydrogen ions’ are coextensive.” Hydrogen ions thus become the primordial acid of the older chemists. The general and characteristic properties of acids are assumed to be the general and characteristic properties of H^+ ions,

and thus the H^+ ions are said to have a sour taste, redden blue litmus, conduct electricity in solutions containing them, behave as univalent radicals, etc. The basicity of an acid is fixed by the number of H^+ ions furnished by the complete ionization of one molecule of the acid. Thus monobasic hydrochloric acid, HCl , furnishes one H^+ ion. $HCl \rightleftharpoons H^+ + Cl^-$; and dibasic sulphuric acid furnishes two H^+ ions: $H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$. Sulphuric acid also furnishes the ions H^+ and HSO_4^- so that it behaves also as a monobasic acid when it forms the so-called "acid sulphates."

Why is the hydrogen ion acidic? The action of a metal, say, zinc, on an acid is usually represented by the equation: $Zn + 2HCl = ZnCl_2 + H_2$; the ionic hypothesis assumes that: $Zn + 2H^+ + 2Cl^- \rightleftharpoons Zn^{2+} + 2Cl^- + H_2$; and that the different atoms have different affinities—electro affinities—for electrical charges. The hydrogen ion is an acid because it holds its charge less tenaciously than do the metal ions. "If it were otherwise," said H. C. Jones, "if the hydrogen ion held its charge as poorly as the average metal, the acids would not be acids." Since the Cl^- ions are but little affected by the change, the last equation reduces to $Zn + 2H^+ = Zn^{2+} + H_2$. The action is thus independent of the negative ion, for it involves little more than a transfer of the positive electric charges from the two hydrogen ions to the zinc, and generally such typical chemical reactions can be represented as the transfer of electrical charges from hydrogen to the metal. When the solution of zinc chloride is concentrated by evaporation, the Zn^{2+} and $2Cl^-$ ions recombine to form zinc chloride.

Bases.—Just as an acid has been defined to be a substance which can furnish hydrogen ions when dissolved in water, so bases, according to the ionic hypothesis, are substances which yield OH^- ions when dissolved in water. The basic properties of bases are due to the OH^- ions, and in this sense it has been said that "there is only one base, and that is the OH^- ions," and that "the two terms 'hydroxyl ions' and 'base' are coextensive." The general and characteristic properties of the bases are supposed to be the general and characteristic properties of the OH^- ions. Thus the OH^- ions are said to have a soapy feel, turn red litmus blue, conduct electricity in a solution containing them, etc. The acidity of a base is fixed by the number of OH^- ions it furnishes on complete ionization of a molecule of the base. Thus, the uniaid bases, like sodium hydroxide, ionize: $NaOH \rightleftharpoons Na^+ + OH^-$; and the biacid bases, like barium hydroxide, ionize: $Ba(OH)_2 \rightleftharpoons Ba^{2+} + 2OH^-$. In a non-ionizing solvent the bases do not furnish hydroxyl ions, and they do not then behave like bases. Hence H. C. Jones defines: "a base is a compound which furnishes hydroxyl ions when dissolved in an ionizing solvent. A compound becomes a base only when it is ionized into hydroxyl ions."

Why is hydrogen an acid in some compounds and not in others? The answer furnished by the ionic hypothesis is that hydrogen separates from the former as an ion, but not so from the latter. Hydrogen chloride in aqueous solution is an acid because it furnishes hydrogen ions, but it is not an acid when dissolved in dry benzene or dry chloroform because it does not furnish hydrogen ions in those solvents. Hence, added H. C. Jones, "an acid is a compound which yields hydrogen ions when dissolved in an ionizing solvent." Dry benzene and dry chloroform are not ionizing solvents. A compound becomes an acid only when it is ionized with hydrogen ions.

Many substances contain hydrogen, and they are not regarded as salts of hydrogen. Methane, CH_4 ; ammonia, NH_3 ; alcohol, $\text{C}_2\text{H}_5\text{OH}$, etc. Again, H_3PO_2 only gives one hydrogen ion per molecule, and the remaining two hydrogen atoms are not ionizable, for they form an essential part of the cation H_2PO_2^+ . Silicic acid is very slightly soluble in water, so that its aqueous solution has no effect on blue litmus. Silicic acid is acid because it forms a salt, sodium silicate, Na_2SiO_3 , which dissolves in water and ionizes: $\text{Na}_2\text{SiO}_3 \rightleftharpoons 2\text{Na}^+ + \text{SiO}_3^{2-}$, when electrolyzed. Neither pure dry hydrogen chloride nor pure dry sulphuric acid is really an acid; they are not dissociated into ions, they do not act on metals, they do not decompose carbonates, nor colour blue litmus.

The chemical activity of different solutions containing equivalent amounts of different acids has been referred to the concentration of the H^+ ions in the solution. The concentration of the H^+ ions depends upon the degree of ionization of the different acids. Hence the relative strengths of the acids can presumably be expressed in terms of the electrical conductivity of equivalent solutions. The speed of a reaction dependent upon an acid is thus connected with the concentration of the H^+ ions. Reverting to the measurements given on p. 119 for hydrochloric, sulphuric, and acetic acids, although the solutions contained equivalent quantities of replaceable hydrogen per litre, acetic acid has but one two-hundredth the activity of hydrochloric acid.

	Hydrochloric acid.	Sulphuric acid.	Acetic acid
Fraction ionized	0.78	0.51	0.004
Relative strength	100	70	0.5

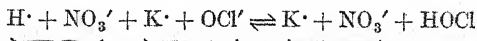
In hydrochloric acid, a greater number of hydrogen ions are ready to react with the metal than with acetic acid, and consequently the available hydrogen in hydrochloric acid is more rapidly exhausted than with acetic acid, where but few ions are in a condition to react with the metal at any moment, and consequently the reaction progresses slowly for a long time; as fast as the available ions are exhausted, new ions are formed by the ionization of the molecule of acetic acid. The total number of hydrogen ions is the same in both cases, but the number in a condition to react with the metal at any moment is very different in all three cases.

Similarly, the strength of a solution containing equivalent quantities of the different bases is referred to the concentration of the OH^- ions in the solution. The strength of a base depends upon the degree of ionization, or on the concentration of the OH^- ions. The strength of a base can thus be determined from the electrical conductivity. In equivalent solutions, bases, like acids, differ very much in strength. The alkalis and alkaline hydroxides are very strong bases, for they are ionized to very nearly the same extent as hydrochloric acid in aqueous solution. Ammonia is a comparatively feeble base. The following numbers represent the relative strengths of a few bases in $\frac{1}{40}$ normal solution on the assumption that the strength of the base is proportional to the electrical conductivity:

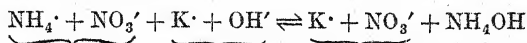
	LiOH	NaOH	KOH	NH_4OH
Relative strength	100	98	98	2

When a highly dissociated acid is mixed with a salt, the two react, forming another acid and salt. The change is reversible, and the reacting

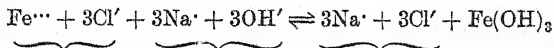
system is then a further illustration of the principle of opposing reactions. For instance, the action of dilute nitric acid on potassium hypochlorite (p. 336), $\text{HNO}_3 + \text{KOCl} \rightleftharpoons \text{KNO}_3 + \text{HOCl}$. If both products are highly ionized, there will be no perceptible change in the system, but in the illustration just cited, hypochlorous acid is but feebly ionized, and, since the H^+ ions of the nitric acid, and the OCl^- ions of the potassium hypochlorite, react to form feebly ionized hypochlorous acid, the result of the reaction in dilute solution is ionized potassium nitrate, and feebly ionized hypochlorous acid:



Similar remarks apply *mutatis mutandis* to the action of a salt on a base, and this explains how feebly ionized ammonium hydroxide is formed in relatively large quantities when highly ionized solutions of potassium hydroxide and ammonium nitrate are mixed together. The reaction proceeds almost to the end:



When the base is insoluble, it will be precipitated and the reaction will proceed to an end quite apart from the degree of ionization of the reacting compounds. This is the case, for example, with ferric, aluminium, zinc, and other hydroxides:



§ 12. The Strengths of Acids and of Bases.

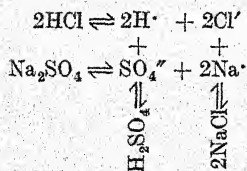
I have no doubt that fixed salts choose one acid rather than another in order that they may coalesce with it in more intimate union.—JOHN MAYOW (1674).

The strength of an acid or base refers to the extent to which the acid or base exhibits acidic or basic properties respectively. The terms "affinity," "avidity," and "activity" are sometimes employed synonymously with "strength," but there are objections to each of these. The term "strength," too, is often used where "concentration" is really meant. "Concentration" refers to the "quantity of matter in unit volume" expressed in, say, grams per litre, or some other convenient form, say, gram-molecules per litre.

The action of sulphuric acid on sodium chloride (p. 271), which results in the formation of hydrochloric acid, seems to prove that sulphuric acid is stronger than hydrochloric acid; again, when hydrochloric acid is added to a solution of silver sulphate, silver chloride is precipitated. The hydrochloric acid expels the sulphuric acid from its combination with silver: $\text{Ag}_2\text{SO}_4 + 2\text{HCl} = 2\text{AgCl} + \text{H}_2\text{SO}_4$, and it seems as if hydrochloric acid is stronger than sulphuric acid. These two conclusions are contradictory and there must therefore be a fallacy in our reasoning. We have wrongly assumed that the two acids were competing for sodium and for silver under similar conditions. This is not the case. When hydrochloric and sulphuric acids compete for the sodium, the hydrochloric acid, being volatile,

escapes from the system as fast as it is formed; while the non-volatile sulphuric acid alone remains behind. Again, when sulphuric and hydrochloric acids are competing for silver, the hydrochloric acid carries the silver away from the sulphuric acid as an insoluble precipitate of silver chloride. Still further, hydrosulphuric acid is notoriously a very feeble acid, and yet it can displace relatively strong acids from combinations with the metals. Thus, it will precipitate lead sulphide from solutions of lead chloride; copper sulphide from solutions of copper sulphate, etc. Here again, the feeble acid does its work by removing the metal from the solution as an insoluble sulphide.

To compare the relative strengths of the acids, and, *mutatis mutandis*, of the bases, it is necessary that the comparison be made under conditions where the reacting acids and the products of the reaction are in the same physical condition—say, all in solution. Thus, if an equivalent of a solution of sodium hydroxide be mixed with an equivalent of a solution of sulphuric and of hydrochloric acids, the two acids can compete for the one base under the same conditions, and hence the stronger acid will be able to unite with more sodium than the weaker acid. In 1803, C. L. Berthollet pointed out that when a neutral salt is dissolved and an acid is added to the solution, the free acid enters into competition with the combined acid, and they both act on the alkali base in the ratio of their respective concentrations as though no combination had existed. It cannot therefore be said that if all the conditions remain equal, one acid displaces another from the base with which it had been united, but the base is shared between the two acids in the ratio of the concentrations and affinities of the respective acids. It is found experimentally that the same result is obtained when equivalent quantities of sodium hydroxide, sulphuric acid, and hydrochloric acid are mixed together as when equivalent quantities of sodium sulphate and hydrochloric acid, or equivalent quantities of sodium chloride and sulphuric acid, are mixed, provided, of course, the whole of the system has been allowed to stand long enough for equilibrium. This fact is represented by the equation: $2\text{HCl} + \text{Na}_2\text{SO}_4 \rightleftharpoons 2\text{NaCl} + \text{H}_2\text{SO}_4$, which, when translated into the language of ions, reads:



In 1803, C. L. Berthollet wrote:

I maintain that whenever several acids act upon one alkaline base, the action of one of the acids does not overpower that of the others so as to form an insulated compound, but each of the acids has a share in the action proportionate to its capacity for saturation and its quantity.

The proportions of a base shared between two acids, or of an acid between two bases, cannot be determined by the ordinary methods of chemical analysis without disturbing the equilibrium of the mixture. The distribution of an acid between two bases, or of a base between two acids, must be determined by physical processes which do not interfere with the

solution. In illustration, the heat of neutralization of sodium hydroxide by sulphuric acid is 31·38 Cals.; and by hydrochloric acid, 27·48 Cals. If, therefore, on mixing hydrochloric acid with sodium sulphate, all the sulphuric acid were displaced by the hydrochloric acid, the thermal effect resulting from the decomposition of the sodium sulphide, and the formation of the sodium chloride would be $27·48 - 31·38 = -3·9$ Cals. After making a small allowance for secondary reactions between sodium sulphate and sulphuric acid, J. Thomsen found that the thermal value of the reaction was $-2·6$ Cals. Hence it follows that $-2·6 \div -3·9$ or about two-thirds of the hydrochloric acid combines with about two-thirds of the base to form sodium chloride; and about one-third of the sulphuric acid combines with the other third of the base to form sodium sulphate. A similar result was obtained with a mixture of sodium chloride and sulphuric acid as with sodium sulphate and hydrochloric acid. Consequently, in the competition of sulphuric and hydrochloric acids for sodium under comparable conditions, the hydrochloric acid can hold twice as much of the base as the sulphuric acid, and consequently, hydrochloric acid is nearly twice as strong as sulphuric acid.

Similar results have been obtained by measuring the specific gravity, index of refraction, absorption of light, etc. The relative strengths of the different acids have also been determined by measuring the effects of the different acids on the speed of hydrolysis of cane sugar, methyl acetate, etc. The actual numbers obtained by the different methods are not always quite the same, possibly because of the different conditions under which the experiments are made. The results obtained by three different methods are shown in Table XX.

TABLE XX.—RELATIVE STRENGTHS OF ACIDS.

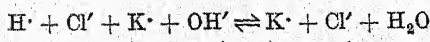
Acid.	Thomsen's thermal process.	Ostwald's specific gravity process.	Molecular conductivity.
Hydrochloric acid	100	98	100·0
Nitric acid	100	100·00	99·6
Hydrobromic acid	89	95·00	100·0
Sulphuric acid	49	66·7	65·1
Phosphoric acid	25	—	7·3
Acetic acid	3	1·23	0·4

§ 13. The Neutralization of Acids and Bases.

The importance for chemistry of the fact that hydrogen and hydroxyl ions cannot remain in the presence of one another uncombined is difficult to overestimate. Could these ions remain separate, then an acid would not neutralize a base, and all salt formation from the process of neutralization of acids and bases would be excluded.—H. C. JONES.

The term "neutral" has been used somewhat vaguely, implying that the substance is neither acidic nor basic. The test for acidity or basicity depended upon the behaviour of the solution towards a solution of litmus. If other indicators are used, the conclusions might be different, because a

substance might appear acidic towards one indicator, and neutral towards another. The ionic hypothesis, as we have seen, refers acidity to the presence of hydrogen ions, and alkalinity to the presence of OH' ions, and the term "neutrality" refers to the case where the concentration of both ions are the same, or both ions are absent. We have seen that water is a constant product of the reaction between the solution of an acid and of a base: $\text{HCl} + \text{KOH} \rightleftharpoons \text{KCl} + \text{H}_2\text{O}$; $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightleftharpoons \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$, etc. When solutions of an acid and base are mixed, the hydrogen and hydroxyl ions of acid and base respectively combine to form water, because water only ionizes to an inconceivably small extent, and the two kinds of ions— H' and OH' —cannot remain in the presence of one another uncombined. Hence, when aqueous solution of acids and bases are mixed together OH' and the H' ions are removed from the solution, and the reaction is almost completed:



What is here stated with respect to hydrochloric acid and potassium hydroxide applies *mutatis mutandis* to any strongly ionized acid and base; and consequently the neutralization of strongly ionized acids and bases involves little more than the formation of water: $\text{H}' + \text{OH}' \rightleftharpoons \text{H}_2\text{O}$, because the other ions present before the acids and bases are mixed remain after the reaction is over. If, however, the water be evaporated from the solution, the ions recombine to form the salt, and the result of the reaction is then correctly symbolized: $\text{HCl} + \text{KOH} \rightleftharpoons \text{KCl} + \text{H}_2\text{O}$. This reaction probably also occurs if very concentrated solutions or solids are mixed, whereas the neutralization of acids and bases in dilute solutions involves the formation of water, not salt molecules.

The heat of neutralization of dilute solutions.—This view is further supported by the fact that with dilute solutions of the strong acids and bases, the thermal value of the process of neutralization—heat of neutralization—is the same. For example,

	LiOH	NaOH	KOH	Ca(OH)_2	Ba(OH)_2
Hydrochloric acid . .	13·7	13·7	13·7	13·8	13·8 Cals.
	HCl	HBr	HI	HNO_3	HIO_3
Sodium hydroxide . .	13·7	13·8	13·7	13·7	13·5 Cals.

Hence neutralization is an isothermal process; the heats of neutralization of dilute solutions of the strong acids and bases do not depend upon the specific nature of the acid or base; and the formation of water in these reactions is accompanied by the evolution of approximately 13·7 Cals. of heat.

The law only describes the thermal effect attending the neutralization of solutions sufficiently diluted to ensure complete ionization of acid, base, and salt; it presupposes that no new electrically neutral molecules are formed. As a corollary, it follows that if two completely ionized salts are mixed, there will be no thermal change, provided the salts are completely ionized before and after the mixing, and no other electrically neutral molecules are formed. The fact that "if two neutral salt solutions at the same temperature are mixed together, no change of temperature

occurs" was discovered by H. Hess in 1841, and is called Hess' law of thermo-neutrality. For example:

Before mixing—		After mixing—	
Calcium nitrate, $\text{Ca}(\text{NO}_3)_2$. .	451	Calcium sulphate, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	642
Potassium sulphate, K_2SO_4 . .	601	Potassium nitrate, KNO_3 . .	409
Thermal value of admixture .		1052	Thermal value of mixture . 1051

The ionic hypothesis indicates clearly the conditions which must be fulfilled before Hess' law of thermo-neutrality is applicable, and, said H. C. Jones, "it would be difficult to find such a successful explanation by any other known hypothesis."

If the acid and base are but partially ionized, the heat of neutralization is not only determined by the heat of formation of water—13·7 Cals.—but it is also determined by the thermal value of the energy required to complete the ionization of acid and base. When a dilute solution of hydrofluoric acid is neutralized by sodium hydroxide, for example, the sodium fluoride formed during the reaction is fully ionized, whereas the hydrofluoric acid at the commencement of the process: $\text{HF} + \text{NaOH} = \text{NaF} + \text{H}_2\text{O} + 16\cdot27$ Cals. is not fully ionized. Hence in addition to the formation of water, there is continuous ionization of hydrofluoric acid *during* the process of neutralization, and the fact that more heat is produced has been assumed to prove that the ionization of the acid is accompanied by the evolution of heat. The heat of neutralization of hypochlorous acid, HOCl , by sodium hydroxide, NaOH is: $\text{HOCl} + \text{NaOH} = \text{NaOCl} + \text{H}_2\text{O} + 9\cdot8$ Cals., a number less than the normal value 13·7 Cals. The salt, NaOCl , and the base, NaOH , are completely ionized; while the acid, HOCl , is but feebly ionized. Hence, it is assumed that the ionization of HOCl is an endothermal process. Similarly, when ammonia is neutralized: $\text{NH}_4\text{OH} + \text{HCl} = \text{NH}_4\text{Cl} + \text{H}_2\text{O} + 12\cdot2$ Cals., it is assumed that the low results are due to the absorption of heat during the ionization of ammonium hydroxide.

Hydrolysis.—It will be remembered (p. 172) that in hydrolysis, a salt reacts with water to form the free base and free acid, or free acid and a basic salt. Hydrolysis is thus a reversion of the process of neutralization. Hydrocyanic acid, HCy , for instance, ionizes: $\text{HCy} \rightleftharpoons \text{H}^+ + \text{Cy}'$. With potassium cyanide, KCy , in aqueous solution, $\text{KCy} \rightleftharpoons \text{K}^+ + \text{Cy}'$. In the latter case, some of the H^+ ions of the water unite with the Cy' ions of the salt to form molecules of hydrocyanic acid, HCy . The equilibrium is disturbed, and more molecules of water ionize: $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}'$. The new H^+ ions combine with more Cy' ions and the process continues until the concentration of the OH' ions becomes large enough to prevent the further ionization of the water. The solution then contains an excess of OH ions, and free hydrocyanic acid, as well as potassium cyanide, and the ions K^+ and Cy' . The free hydrocyanic acid can be recognized by its smell; and the OH' ions can be recognized by the alkalinity of the solution.

The ionic hypothesis in analytical chemistry.—The language of the ionic hypothesis has penetrated into analytical chemistry—particularly qualitative analysis—so that the standard tests for the metals and the acid radicals are described as tests for the respective ions. Many, however, doubt if anything will be gained by describing the facts of an eminently

practical art in the language of a hypothetical doctrine. Be that as it may, since both the chlorides and hydrochloric acid are supposed to furnish Cl' ions on ionization, it is assumed that the test for hydrochloric acid or a chloride is a test for chloride ions. The silver nitrate used in making the test is also supposed to be ionized. Consequently, when a solution of silver nitrate is added to a solution of sodium chloride, the solution is supposed to momentarily contain $\text{Ag}' + \text{NO}_3' + \text{Na}' + \text{Cl}'$ ions, but since silver chloride is but slightly soluble in water, the little which is formed by contact of silver and chlorine ions precipitates at once. Again, silver chlorate is soluble in water, and accordingly, when silver nitrate is added to a solution of potassium chlorate there is no precipitation and the solution contains four kinds of ions $\text{Ag}' + \text{NO}_3' + \text{K}' + \text{ClO}_3'$. Hence, silver nitrate is a test for chloride but not for chlorate ions.

If potassium cyanide, KCy , in aqueous solution be added to a solution of silver nitrate, AgNO_3 , a precipitate of silver cyanide is formed: $\text{Ag}' + \text{NO}_3' + \text{K}' + \text{Cy}' = \text{AgCy} + \text{K}' + \text{NO}_3'$. If an excess of potassium cyanide be added, the precipitate redissolves, and it can now be shown that the solution no longer contains the equivalent of Ag' ions in appreciable quantities, since (1) sodium chloride gives no precipitation of silver chloride; (2) on electrolysis, silver collects about the anode, not the cathode, as when a solution of silver nitrate is electrolyzed; and (3) a crystalline compound KAgCy_2 is obtained when the solution is concentrated. It is assumed, therefore, that the solution of silver cyanide in potassium cyanide ionizes thus: $\text{KAgCy}_2 \rightleftharpoons \text{K}' + \text{AgCy}_2$.

Questions.

1. What is Ostwald's dilution formula for weak electrolytes? Deduce it theoretically. Explain clearly how the constant is experimentally obtained.—*St. Andrews Univ.*
2. What is meant by "electrolysis"? Illustrate your answer by reference to the electrolysis of aqueous solutions of metallic salts.—*St. Andrews Univ.*
3. Write the following equations according to the ionic hypothesis: (a) Calcium hydroxide and hydrochloric acid; (b) Barium chloride, and sulphuric acid.—*Univ. of Pennsylvania, U.S.A.*
4. Explain in terms of the electrolytic dissociation theory: (a) The interaction of a solution of copper sulphate with metallic iron. (b) The electrolysis of copper chloride; (c) The interaction of a solution of silver nitrate and hydrochloric acid; (d) Neutralization; (e) The behaviour of hydrogen chloride in toluene solution and in water solution.—*Princeton Univ., U.S.A.*
5. In terms of the ionic theory, what is a strong acid?—a strong base?—a weak acid?—a weak base?
- In these terms classify the following:— HCl , NH_4OH , KOH , H_2S , H_2CO_3 , $\text{Al}(\text{OH})_3$.—*Amherst Coll., U.S.A.*
6. Define "acids," "bases," and "salts" in terms of (a) their properties; (b) their composition; and (c) the ionic theory.
7. Give a brief history of the word "acid" with an account of its application in the present day. What do you understand by the term "ortho-acid"?—*London Univ.*
8. Give an account of the ionic theory of solution, stating clearly the experimental facts on which it is based.—*Aberystwyth Univ.*
9. A current passes simultaneously through acidulated water, a solution of copper sulphate, CuSO_4 , and molten silver chloride. What substances are produced in each cell, and how many grams of each in the time that 10 cubic centimetres of hydrogen are liberated from water ($\text{Cu} = 63$, $\text{Ag} = 108$)?—*New Zealand Univ.*
10. Explain the italicized words: "Alumina like boric oxide is *amphoteric*."

The *acidic and basic qualities* depend on the presence of both OH' and H' ions, and both ions cannot be present in very great *concentration* in the same solution."

11. Explain the meaning of the italicized terms: "To make the *qualitative analysis* decisive, add a *strong solution* of ammonia to the *precipitate*; if it dissolves, *neutralize* the ammonia with nitric acid, etc."

12. Describe from the ionic point of view what happens when—(a) 1 gram of sodium chloride is dissolved in 1 litre of water: (b) the solution is evaporated down until it occupies 20 c.c.: (c) it is saturated with gaseous hydrochloric acid so that the solid salt separates.—*Calcutta Univ.*

13. Explain what is meant by the "avidity" of an acid. Discuss its theoretical bearings and show how "avidity" may be experimentally determined.—*London Univ.*

14. J. H. Gladstone and W. Hibbert (1889) passed the same current of electricity through solutions of zinc and silver salts, and obtained quantities of these metals in the ratio $Zn : Ag = 1 : 3.298$. If the equivalent of silver is 107.88, what is that of zinc?

15. What do you understand by the term "relative migration velocity"? In what way is the relative migration velocity connected with the molecular conductivity of an electrolyte? Describe an experiment by which such a velocity can be determined.—*Board of Educ.*

16. One and the same electric current is simultaneously passed through solutions (a) hydrochloric acid, (b) copper sulphate, (c) silver sulphate. Draw a diagram of the apparatus required, and indicate what products would be obtained in the case of each solution. Calculate also the weights of these products which would be formed during the time that 1000 c.c. of dry hydrogen measured at 0° and 760 mm. were collected from the hydrochloric acid solution. The formation of peroxides is to be neglected. (Atomic weights, $H = 1$, $Cl = 35.5$, $Cu = 63$, $Ag = 108$, $S = 32$, $O = 16$.)—*London Univ.*

17. How is the following phenomenon explained? When dilute solutions of any strong acid and base are mixed together in equivalent proportions, approximately the same amount of heat (13,700 Cals.) is given out, although the heats of formation of the different salts formed have very different values.—*Madras Univ.*

18. What is meant by the transport number of an ion? Explain how it can be measured in practice, and sketch the apparatus employed. Calculate the transport number of the cupric ion from the following data: In the neighbourhood of the cathode, a given volume of a solution of copper sulphate gave before electrolysis 0.6765 grm. of copper oxide, CuO , and after electrolysis, 0.5118 grm. of copper oxide, CuO . During the electrolysis, 0.2045 grm. of copper had been deposited on the cathode.—*Panjab Univ.*

CHAPTER XVII

THE ALKALINE EARTHS

§ 1. Calcium Carbonate.

THE early chemists reserved the term *earth* for those substances which were insoluble in water and which did not undergo alteration when calcined at a high temperature—*e.g.* alumina, silica, magnesia, lime, etc. The earths—lime and magnesia—which were related to the alkalies by giving an alkaline reaction and neutralizing acids, were termed *alkaline earths*. Baryta and strontia were afterwards included among the alkaline earths.

Several different minerals occur in nature which have been called different names, although analysis shows that they are all more or less impure forms of one chemical substance—calcium carbonate. These different forms of calcium carbonate may be classed under three heads, although in reality there are but two crystalline forms or mineral species.

1. **Calcium carbonate in rhombic crystals.**—This variety generally occurs in needle-like crystals, and is named *aragonite* after Aragon in Spain. If calcium carbonate be prepared in solutions at temperatures exceeding 30° , crystals corresponding with aragonite are formed, and if at temperatures below 30° , crystals of calcite are formed. Hence, aragonite at temperatures below 30° is in a metastable condition. The fundamental form of crystals of aragonite is illustrated by the outline drawings, Fig. 139, and although there are a great many derived shapes, all are built on the same geometrical plan determined by the inclination of the boundary faces with respect to the axes of the crystal.

2. **Calcium carbonate in trigonal crystals.**—This form of calcium carbonate occurs in more or less well-defined crystals modelled after a rhombohedron, Fig. 139, but exhibiting a great variety of derived shapes which have received special names—"dog's-tooth spar," "nail-headed spar," etc. R. J. Haüy called calcite the *proteus* among minerals because of its presenting what appeared to be a chaotic number of unrelated appearances, yet he also showed that, however diverse the forms of the crystals, they are all related to a primitive geometrical plan which is determined by the inclination of the boundary faces with the crystal axes. All the different forms are subordinate to the primitive type. When transparent and colourless, trigonal calcium carbonate is called *Iceland spar*; and if opaque and clouded, *calcite* or *calcspar*. A compact fibrous variety with a satin-like lustre is called "satin spar." *Marble* is made up of minute crystals of calcite. *Onyx* is a variety which is streaked and coloured by

associated impurities. Aragonite has a specific gravity about 2.93, calcite 2.72; the heat of formation of aragonite is greater than that of

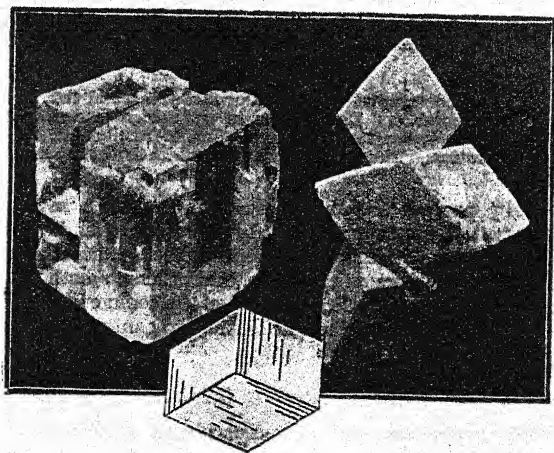


FIG. 139.—Crystals of Calcite from Fontainebleau.

calcite, being respectively 21.33 and 21.03 Cals.; and the specific heat of aragonite is 0.1899, while that of calcite is 0.1887.

3. Calcium carbonate not markedly crystalline.—Chalk and limestone usually occur in large masses sometimes extending over large tracts

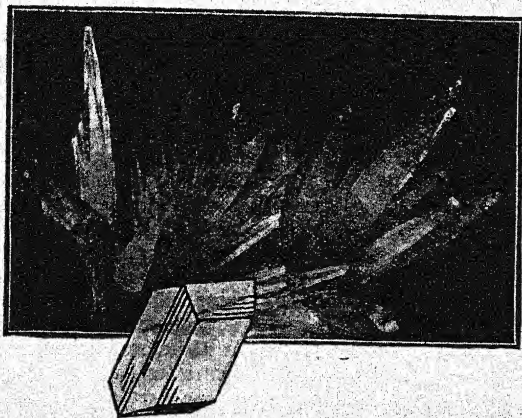


FIG. 140.—Crystals of Aragonite from Cumberland.

of country. This form of calcium carbonate is relatively impure, for it contains more or less magnesium carbonate, clay, and silica. Marl is a

mixture of limestone and clay. Egg-shells, sea-shells, pearls, corals, contain a large percentage of calcium carbonate. The microscope shows that *chalk* consists largely of the shells of minute organisms. It is inferred that these organisms once lived in the sea because similar shells are dredged from the bottoms of the oceans to-day. Chalk is used in the manufacture of whiting.

The chalk is ground to powder, stirred up with water, and the slurry allowed to stand for a few minutes. The matter in suspension is decanted from the sandy sediment. The water is removed by further settling and decantation, and also by evaporation. The result is whiting, which is therefore chalk minus the sandy grit.

H. C. Sorby proved that in calcareous organisms a new variety of calcium carbonate is not in question. In every case examined, the shells contained either calcite or aragonite. Some animal species secrete calcium carbonate as calcite, others as aragonite, and others as both. The pearl, for instance, is mainly aragonite; the shell of the common whelk is mainly calcite; and the inner shell of the cuttle-fish is aragonite, the outer portion is calcite.

§ 2. The Action of Heat on Calcium Carbonate.

If any one of these forms of calcium carbonate be heated by means of a gas blow-pipe flame in a hard glass or quartz test tube, fitted with a gas delivery tube as illustrated in Fig. 47, a gas called carbon dioxide, symbolized CO_2 , collects in the gas jar, and a residue called calcium oxide or quicklime, symbolized CaO , remains in the test tube. The gas will be studied later. It is sufficient to state here that the moist gas reddens blue litmus, and unites with bases to form salts called carbonates. If, say, 1.00 gram of pure calcium carbonate be thoroughly calcined in a crucible, 0.56 gram of calcium oxide is obtained, and 0.44 gram of carbon dioxide is expelled as gas. If a current of carbon dioxide be passed over calcium oxide, calcium carbonate, symbolized CaCO_3 , is formed:

$\text{CaO} + \text{CO}_2 = \text{CaCO}_3$. A certain amount of heat is developed during this latter reaction. Calcium carbonate is undoubtedly a chemical compound of calcium oxide and carbon dioxide.

If calcium carbonate be heated in a closed vessel, at different temperatures, when the system is in equilibrium, the state of the system will be represented by a point on the curve, Fig. 141, where the pressures in the closed vessel are plotted at different temperatures. There are three phases— CaO , CaCO_3 , and CO_2 ; and two components— CaO , and CO_2 . The system is therefore univariant, meaning that the concentration—that is the pressure—of the gas has one fixed definite value for each temperature. This constant pressure is called the dissociation pressure. The dissociation pressure of calcium carbonate is analogous, in

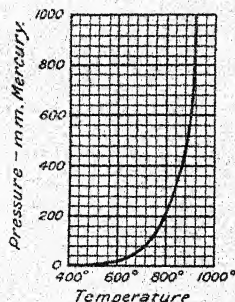


FIG. 141.—Effect of Temperature on the Dissociation Pressure of Calcium Carbonate.

many ways, with the vapour pressure of a liquid in a closed space. Gaseous molecules of carbon dioxide are continually leaving the disso-

ciating carbonate, and carbon dioxide is continually recombining with calcium oxide. When the speeds of the two reactions are the same, the system is in equilibrium. The carbonate can only dissociate completely into calcium oxide and carbon dioxide if the latter be continuously removed from the reacting system. If the pressure, at any temperature, exceeds the limit indicated by the curve, calcium carbonate will be formed until the pressure of the gas attains the fixed value, constant for the given temperature; and conversely, if the pressure be less than that described by the curve, calcium carbonate will dissociate until the required pressure is attained. If the pressure of the gas be great enough, the calcium carbonate may be fused with no appreciable dissociation. On solidification, the mass has a crystalline structure like marble. This shows how, in nature, marble can be associated with igneous rocks whose temperature at certain geological epochs must have greatly exceeded the dissociation temperature of the carbonate. Under an enormous pressure, the carbonate would melt and subsequently crystallize without decomposition. J. Hall verified this experimentally in 1805.

According to the law indicated in connection with the iron-steam reaction, the velocity of the dissociation of calcium carbonate will be proportional to the concentration of the carbonate, and equal to the product of the concentration C of the carbonate and the affinity constant k ; the velocity of formation of the calcium carbonate will similarly be equal to the product of the concentration of the calcium oxide, C_1 , and of the carbon dioxide, C_2 , with the affinity constant k' of that reaction. Consequently, if these two velocities are the same, the system will be in equilibrium, and $kC = k'C_1C_2$. The concentration, that is, the number of molecules of carbon dioxide per litre, is proportional to the partial pressure p of that gas. It is assumed that all solids, calcium carbonate and calcium oxide, exert a small vapour pressure which is generally too small to come within the range of the instruments at present available for such measurements. This vapour pressure is further assumed to be constant at any given temperature, like the vapour pressures of substances which are accessible to measurement; and is also unaffected by the quantity of solid present. The concentrations C and C_1 are constant, p. 121, and consequently, at any given temperature, $p = \text{constant}$. This means that the dissociation pressure of the carbon dioxide is constant, and independent of the extent to which the calcium carbonate has dissociated, provided all the carbonate has not dissociated. This same conclusion was obtained in applying the phase rule. The principles here stated also apply to the dissociation of mercuric oxide, of hydrates, and of barium peroxide, as previously indicated.

According to A. Colson, dry silver carbonate, like calcium carbonate, has a definite dissociation pressure at a given temperature, but the reaction is not reversible because the silver oxide and carbon dioxide do not recombine on cooling. Lead carbonate behaves similarly. If, however, a trace of moisture be present, the reactions are reversible. Presumably the silver and lead oxides undergo some change soon after they are formed which hinders the reverse reaction. C. W. Scheele also observed, in 1777, that dried calcium oxide does not absorb dry carbon dioxide.

§ 3. Calcium Oxide and Calcium Hydroxide.

Calcium oxide is nearly always made by calcining the carbonate—marble, Iceland spar, limestone. The residue is variously styled “quicklime,” “live lime,” “burnt lime,” or “caustic lime”—from the Greek *καυστικός* (kaustikos), burnt. Calcium oxide, when pure, is a white amorphous powder. If heated intensely, say in the oxyhydrogen blowpipe, it becomes incandescent, p. 124. In the electric furnace, calcium oxide can be melted at a temperature of about 2570° , and at a still higher temperature it can be boiled. When a few drops of water are allowed to fall on a (cold) lump of freshly “burned” calcium oxide, which has not been burned at too high a temperature, a hissing noise is produced, and clouds of steam arise. Much heat is developed, and the lump of calcium oxide disintegrates into a fine powder called “slaked lime,” or calcium hydroxide. The term “lime” is often applied to quicklime calcium oxide, and also to slaked lime, calcium hydroxide. Moist lime turns red litmus blue. If calcium oxide be treated with water in a weighed dish, and the result of the reaction be thoroughly dried by heating to 150° , 5.6 grams of calcium oxide furnish 7.4 grams of calcium hydroxide, showing that 56 grams of calcium oxide unite with 18 grams of water to form calcium hydroxide— $\text{CaO} \cdot \text{H}_2\text{O}$, or as is more commonly symbolized, $\text{Ca}(\text{OH})_2$, and the reaction is written: $\text{CaO} + \text{H}_2\text{O} = \text{Ca}(\text{OH})_2$. When calcined at 100° , calcium hydroxide loses no water; at 400° , 30 per cent. of the possible water is expelled; and at 450° , nearly all the water may be driven off. When a considerable amount of calcium hydroxide is suspended in water, the mixture is called “milk of lime,” and if the mixture be allowed to stand, a clear liquid collects above the white sediment. The solution—called “lime water”—has a characteristic taste, and it turns red litmus blue. At ordinary temperatures, 10°, 100 c.c. of water dissolves 0.17 gram of the hydroxide, $\text{Ca}(\text{OH})_2$. The solubility diminishes with rise of temperature, for instance, the solubility decreases from 0.185 gram of $\text{Ca}(\text{OH})_2$ at 0° , to 0.128 at 50° , and 0.077 at 100° per 100 grams of water.

When lime water is exposed to the air it soon becomes covered with a film of calcium carbonate, owing to the absorption of carbon dioxide from the atmosphere: $\text{Ca}(\text{OH})_2 + \text{CO}_2 = \text{CaCO}_3 + \text{H}_2\text{O}$.

J. Kolb found, in 1867, that dried calcium hydroxide does not absorb dry carbon dioxide. Both calcium oxide and hydroxide absorb the gas only if a trace of moisture be present. This is usually explained by assuming that carbon dioxide can act on these substances, only when it is converted into carbonic acid, H_2CO_3 , by combination with water. The reaction then proceeds: $\text{CaO} + \text{H}_2\text{CO}_3 = \text{CaCO}_3 + \text{H}_2\text{O}$, and the liberated water reforms carbonic acid. Thus the water acts like a catalytic agent in a long series of cyclic transformations.

Mortar is a thick paste made by mixing slaked lime with sand and water. Mortar sets or hardens on exposure to the air owing to the loss of water, and the absorption of carbon dioxide from the atmosphere. It appears as if a colloidal calcium hydroxide is formed when lime is wetted, and this, on drying, “sets” by a process analogous with the drying of glue. The formation of calcium carbonate is quite a late stage in the setting. In the mortars of some of the buildings of the ancient Romans, crusts of minute

calcium carbonate crystals protect the inner cores of calcium hydroxide. An exposure of 2000 years has not been sufficient to convert more than thin superficial layers of the colloidal hydroxide into carbonate. A little of the colloidal hydroxide seems to pass into the crystalline condition. The sand makes the mass more porous and facilitates the rapid absorption of carbon dioxide, and it prevents undue shrinkage during setting. The sand is scarcely altered during the action. If mortar be placed between bricks or stones, part of the water is absorbed by the bricks and part is lost by evaporation; and the mortar, when set, holds the bricks or stones firmly in place. If mortar is to be used for plastering walls, it is mixed with hair which makes the wet mortar stick better. The setting of mortar is said to be accelerated if a trace of aluminium chloride is present.

Uses of lime.—Lime is used in preparing mortars and cements, as a flux in metallurgy, in the manufacture of glass, bleaching powder, calcium carbide, in purifying illuminating gas, in removing hair from hides before tanning, in drying gases, as a disinfectant, in marling soil, as a neutralizing agent for acids in chemical industries, etc.

Manufacture of quicklime.—Limestone is frequently "burned" in a cavity cut in a hill-side, or in a kiln made from bricks or blocks of limestone. The kiln is loosely filled with limestone, or alternate layers of limestone and fuel. A fire is built at the base of the stack, and when the burning is complete, the fire is allowed to die out, and the lime is removed. In the more recent "shaft kilns," producer gas is led into the shaft near the base, and the hot products of combustion pass up the stack and decompose the limestone. The kiln is charged at the top, and the burned lime is raked out through openings at the bottom of the stack. The process is continuous—lime is charged in at the top as fast as it is removed at the bottom.

Hydraulic cements.—Calcareous marls or mixtures of limestone, clay, and sand in the right proportions are fed into a long slanting slowly rotating cylindrical furnace, and there heated until the mixture begins to sinter. The "clinker" so obtained is ground to powder, and the product called "cement," because if it be mixed with water it sets to a hard stone-like mass, even if exposed to the continued action of water. The cement is consequently used for work under water—bridge-piers, etc.—under conditions where mortar would disintegrate and soften. Several varieties are on the market under various names—Portland cement, hydraulic cement, etc. The ground clinker appears to consist mainly of a solid solution of lime, silica, and alumina which, for convenience, is sometimes called alite. The setting of the wetted cement appears to be the joint effect of several different actions not yet clearly understood. The main reaction is probably due to the formation of colloidal calcium aluminosilicates when "alite" is wetted. These decompose, forming crystals of tricalcium aluminate and a number of different substances in a colloidal condition. The gradual desiccation of the colloids leads to the gradual hardening of the mass; the desiccation appears to proceed even when the cement is immersed in water.

§ 4. Strontium and Barium Oxides and Hydroxides.

Strontium carbonate is found in nature in the mineral *strontianite*, SrCO_3 ; and barium carbonate in the mineral *witherite*, BaCO_3 . Barium

carbonate decomposes at about 1842° , strontium carbonate at 1150° , and calcium carbonate at 825° . When strontium carbonate is calcined, strontia, that is, strontium oxide, SrO , is formed; witherite furnishes baryta, *i.e.* barium oxide, BaO . Strontia is made on a large scale by heating the carbonate in superheated steam; carbon dioxide is evolved, and strontium hydroxide, $\text{Sr}(\text{OH})_2$, is formed: $\text{SrCO}_3 + \text{H}_2\text{O} = \text{Sr}(\text{OH})_2 + \text{CO}_2$. The strontium hydroxide on ignition furnishes strontium oxide: $\text{Sr}(\text{OH})_2 = \text{SrO} + \text{H}_2\text{O}$. Barium carbonate requires so high a temperature for its decomposition that the raw material is mixed with lampblack or tar before calcination. The carbon burns off, and the carbonate, at the same time, is decomposed at a much lower temperature: $\text{BaCO}_3 + \text{C} = \text{BaO} + 2\text{CO}$.

Like calcium oxide, both strontia and baryta slake in contact with water with the evolution of much heat. In the case of baryta, the heat is so great that if but little water be used, the mass may become visibly red hot. The heats of formation of the different hydroxides are: $\text{Ca}(\text{OH})_2$, 16.25 Cals.; $\text{Sr}(\text{OH})_2$, 17.70 Cals.; and $\text{Ba}(\text{OH})_2$, 22.26 Cals. Barium hydroxide is usually made by heating a mixture of *barytes*—barium sulphate, BaSO_4 —with powdered coke, or coal. Crude barium sulphide is formed: $\text{BaSO}_4 + 4\text{C} = 4\text{CO} + \text{BaS}$. The latter is then heated in a stream of carbon dioxide, and thus converted into barium carbonate: $\text{BaS} + \text{CO}_2 + \text{H}_2\text{O} = \text{BaCO}_3 + \text{H}_2\text{S}$. Barium carbonate is converted into the hydroxide by heating it in superheated steam as just indicated for strontium carbonate. Strontium hydroxide is formed in a similar manner from the mineral *celestine*—strontium sulphate, SrSO_4 . Strontium hydroxide is used in the manufacture of sugar. Strontium hydroxide is more soluble in water than calcium hydroxide, and barium hydroxide is more soluble than strontium hydroxide. The solubilities of the three hydroxides in grams per 100 c.c. of water are:

	0°	50°	100°
Calcium hydroxide . . .	0.18	0.13	0.08
Strontium hydroxide . .	0.41	2.5	21.7
Barium hydroxide . . .	1.67	13.12	101.4

Unlike calcium hydroxide, barium hydroxide can be fused without decomposition. Aqueous solutions of both barium and strontium hydroxides deposit isomorphous crystals with eight molecules of water—*e.g.* $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$. Like strontium hydroxide, barium hydroxide is dehydrated when heated high enough, forming barium oxide. The metals of the alkaline earth metals do not form peroxides by direct oxidation as is the case with the alkali metals, but the peroxides are made by oxidizing the oxides; for instance, if barium oxide, BaO , be calcined in air, it forms barium peroxide, BaO_2 , as previously indicated. Barium, calcium, and strontium peroxides are made by adding hydrogen peroxide to solutions of the corresponding hydroxides. The peroxides crystallize out with eight molecules of water—*e.g.* $\text{CaO}_2 \cdot 8\text{H}_2\text{O}$. The hydrated peroxides lose water at about 130° , forming the corresponding anhydrous peroxides. When more strongly heated they decompose, giving oxygen: $2\text{CaO}_2 = 2\text{CaO} + \text{O}_2$.

§ 5. Metallic Calcium, Strontium, and Barium.

Impure calcium was made by H. Davy in 1808, and the pure metal was made by H. Moissan, in 1898, by reducing calcium iodide with sodium.

The metal is now made by electrolysis of the fused chloride, CaCl_2 . In G. O. Seward and F. von Kügelgen's process (1908), the anode is a graphite crucible, and the cathode a rod of iron which dips in the fused chloride, Fig. 142. When the current passes, metallic calcium collects at the lower end of the cathode. Calcium chloride fuses at a lower temperature than metallic calcium, and the temperature is so regulated that the calcium solidifies on the cathode. An irregular rod of metallic calcium is made by slowly raising the cathode. The end of the calcium rod, dipping in the fused chloride, then forms the lower end of the cathode. The rod of metallic calcium dipping in the bath is also cooled by an annular tube—not shown in the diagram—through which cold water flows. Metallic strontium and barium, more or less impure, were first prepared by a similar process to that used for metallic calcium by H. Davy about 1808.

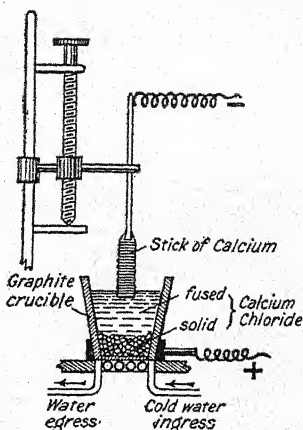


FIG. 142.—The Electrolytic Process for Calcium (Diagrammatic).

Properties.—Calcium is a silver white lustrous metal. It tarnishes slowly in air, and when heated in an atmosphere of hydrogen, it forms calcium hydride, CaH_2 ; and in an atmosphere of nitrogen, calcium nitride, Ca_3N_2 ; in air, calcium oxide, CaO ; and with chlorine, calcium chloride, CaCl_2 . The interaction of water and calcium is rapid, but not violent. Hydrogen and calcium hydroxide are produced: $\text{Ca} + 2\text{H}_2\text{O} = \text{Ca(OH)}_2 + \text{H}_2$. A crust of calcium hydroxide, Ca(OH)_2 , forms on the surface of the metal, and slows down the reaction. If an acid be present, the reaction is rather violent. The properties of strontium and barium resemble those of metallic calcium, all being silvery white metals.

Atomic Weights.—Analyses of calcium chloride, calcium carbonate, etc., give numbers corresponding with a combined weight between 40.03 and 40.23; the best representative value is taken to be 40.09 (oxygen = 16). Several different lines of evidence, to be discussed later, show that this number is not far from the atomic weight. According to Dulong and Petit's rule, for example, the quotient of 6.4 divided by the specific heat of a solid element is a close approximation to the atomic weight of that element. The specific heat of calcium is 0.152. Hence $6.4 \div 0.152 = 42$ approximates to the atomic weight of calcium. Barium and strontium respectively furnish the numbers 137.37 (extremes 137.10 and 137.38) and 87.63 (extremes 87.37 and 87.68) for the atomic weights.

Occurrence of these elements in nature.—The free elements do not occur in nature. Calcium compounds are rather abundant. The occurrence of calcium as carbonate has already been discussed. Calcium also occurs as sulphate in *gypsum* or *selenite*, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$; as calcium fluoride, CaF_2 , in *fluorspar*; as calcium phosphate in *phosphorite*, etc. (*q.v.*). The carbonate and sulphate are common in spring and river water. Calcium compounds occur in all animal and vegetable organisms. Bones contain

a large proportion of combined calcium. The chief natural compounds of barium and strontium have already been indicated—*strontianite*, SrCO_3 ; and *celestine*, SrSO_4 ; *heavy spar* or *barytes*, BaSO_4 ; and *witherite*, BaCO_3 . It is worthy of note that if any two of these elements are found in the same mineral, the pairs will probably be barium and strontium, or strontium and calcium, for barium and calcium do not usually pair together in the absence of strontium—*e.g.* strontium is common in witherite, strontium in aragonite, and calcium in celestine.

History.—The name "calcium" is derived from the Latin *calx*, for lime. Lime was not distinguished from the other earths by the early chemists, but towards the middle of the eighteenth century lime was recognized as a distinct earth.

"Strontium" is named after Strontian, a village in Argyllshire (Scotland), where strontianite was first found (1787). The mineral was first confounded with barium carbonate. In 1790 A. Crawford suggested that strontianite contained a peculiar earth, and based his opinion on the experiments of W. Cruickshank. The suggestion was confirmed by T. C. Hope (1792), and by M. H. Klaproth (1793).

V. Casciorolus, in 1602, noticed that when heavy spar was calcined with combustible matters, the product became phosphorescent in the dark. He called the stone *lapis solis*, and later, it was called "Bolognian," or Bononian phosphorus. The heavy spar which furnished *lapis solis* was at first believed to be a peculiar kind of gypsum. K. W. Scheele (1774) found that the mineral contained a new earth which gave a sulphate insoluble in water. G. de Morveau called the earth "barote"—from the Greek *βαρύς* (*barus*), heavy—and Lavoisier later altered the word to "baryta," the name now used for this earth.

§ 6. The Relations between Calcium, Strontium, and Barium.

The elements of the alkaline earths—calcium, barium, and strontium—exhibit a close kinship and display a gradation in their properties from member to member as the atomic weight increases in passing from calcium to barium. The elements become more active chemically as their atomic weight increases; the specific gravities of the elements and compounds increase; the basic properties and solubilities of the hydroxides increase; and the solubilities of the halides, nitrates, sulphates, and chromates decrease. The physical properties of the elements are indicated in Table XXI, next page.

The metals are fairly stable in air; they quickly tarnish in ordinary air; and when heated, they burn to the monoxide; they are all bivalent; they combine with water with the evolution of hydrogen at ordinary temperatures; and form soluble oxides of the type $\text{R}'\text{O}$, hydroxides of the type $\text{R}'(\text{OH})_2$, and peroxides of the type $\text{R}'\text{O}_2$. The salts are discussed under "Chlorides," "Sulphates," "Nitrates," "Sulphides," etc. The normal carbonates are but sparingly soluble in water; the unstable acid carbonates are more soluble. The low solubility of the carbonates, chromates, and sulphates is utilized in analytical work. All three carbonates are precipitated when ammonium carbonate is added to solutions of their salts. Radium, in this family of elements, will be discussed later.

TABLE XXI.—PHYSICAL PROPERTIES OF THE ALKALINE EARTH METALS.

	Calcium.	Strontium.	Barium.
Atomic weight	40.09	87.63	137.37
Specific gravity	1.52	2.55	3.75
Atomic volume	26.4	35.15	36.6
Melting point	780°	800°	850°
Subliming point	720°	770°	820°
Specific heat	0.152	—	0.068
Flame coloration	Brick-red	Crimson	Green
Heat of formation of monoxide	131.3 Cals.	130.98 Cals.	130.38 Cals.
„ of dioxide from monoxide	4.11 Cals.	13.07 Cals.	18.36 Cals.
„ of hydroxide	16.3 Cals.	17.7 Cals.	22.3 Cals.
Carbonate decomposes	825°	1150°	1842°
Hydroxide decomposes	547°	778°	998°
Solubility of hydroxides (50°)	0.13	2.5	13.1
„ sulphates (10°)	0.19	0.10	0.0002
„ chlorides (50°)	126	72	44

Questions.

1. What is lime? How is it obtained? What takes place (a) when lime is mixed with water; (b) when it is heated **strongly** with sand; (c) when it is exposed to carbon dioxide gas? Give equations.—*Aberdeen Univ.*

2. Distinguish between quicklime, slaked lime, lime water, and milk of lime. What do they absorb from the air?—*Sheffield Scientific School, U.S.A.*

3. Give the names and formulæ of the four principal minerals containing calcium. How is metallic calcium prepared? By what reactions would it be possible to prepare from the metal specimens of (a) calcium dioxide, (b) calcium carbonate, (c) bleaching powder?—*London Univ.*

4. Marignac found that 5 grams of strontium chloride containing six molecules of water of crystallization gave 3.442 grams of strontium sulphate. Calculate the equivalent of strontium. (H = 1, O = 16, S = 32, Cl = 35.5).—*London Univ.*

5. If the heats of formation of the following calcium compounds are nearly:

Iodide	Oxide	Bromide	Chloride
118.6	122	152	170 Cals.

explain how chlorine will displace bromine; bromine, oxygen; and oxygen, iodine from combination with calcium.

6. Describe clearly how barium oxide can be employed for the preparation of oxygen. Barium oxide can be made by heating witherite and nitric acid together. What weights of witherite and nitric acid are required to prepare a kilogram of barium oxide? Assume that nitric acid of sp. gr. 1.40 contains 65.3 per cent. of HNO_3 .—*Bombay Univ.*

7. How does calcium occur in nature? What are its principal compounds in everyday use, and how are they made from naturally occurring compounds? How is metallic calcium made? Give the formulæ of the typical salts of calcium, and of the ions formed on solution of these salts in water.—*Sydney Univ.*

CHAPTER XVIII

BERYLLIUM, MAGNESIUM, ZINC, CADMIUM, AND MERCURY

§ 1. Beryllium and Magnesium.

History of beryllium.—While analyzing beryl, in 1797, L. N. Vauquelin found that a precipitate which he thought to be aluminium hydroxide, dissolved like aluminium hydroxide in potassium hydroxide; but unlike aluminium hydroxide, the solution furnished a white precipitate when boiled for some time. Unlike aluminium hydroxide, too, the precipitate was soluble in ammonium carbonate, and behaved in many other ways differently from aluminium hydroxide. Hence L. N. Vauquelin announced the discovery of a new earth—"la terre du Beril," in 1798. The editors of the *Annales de Chimie*, in which the discovery was announced, proposed the name "glucine"—from the Greek γλυκύς (glucus), sweet—because many of the salts of Vauquelin's "le terre du Béril" had a sweet taste. Since other salts possessed the same property, the term *beryllia*—derived from the name of the mineral—has almost displaced the term "glucina" from recent chemical literature. F. Wöhler isolated the metal beryllium in 1828 by the action of potassium on beryllium chloride.

Occurrence and extraction of beryllium.—The mineral *beryl*, $3\text{BeO} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$, is the principal source of beryllium. The beryls include the gem-stones: *emerald* (pale green), and *agua marine* (pale blue). Beryllium salts can be obtained from beryl by fusing the mineral with sodium carbonate; digesting the resulting mass with hydrochloric acid; evaporating the solution to dryness to separate the silica in an insoluble condition; extracting the soluble matters with dilute hydrochloric acid; precipitating a mixture of aluminium and beryllium hydroxides with ammonia; dissolving the precipitate in potassium hydroxide; and boiling the solution so as to precipitate the beryllium hydroxide. The precipitate is re-dissolved in acid and re-precipitated from the potash solution a number of times to ensure its freedom from alumina. The hydroxide can then be converted into various salts by dissolving it in the proper acid—hydrochloric acid for beryllium chloride, etc.

History of magnesium.—In 1695 N. Grew published a pamphlet describing a peculiar salt found in the mineral springs at Epsom. The medicinal properties of this salt attracted some attention; in England the salt was called "Epsom salt," and on the Continent, "sal anglicum." Magnesia alba (a basic magnesium carbonate) came into commerce from Rome about 1700; the term "magnesia alba" was applied to the earth owing to some fanciful contrast with "magnesia nigra," the term then used for black oxide of manganese. In 1755 J. Black clearly distinguished

between magnesia and lime by showing that magnesia furnished a soluble sulphate, and lime a sparingly soluble sulphate. When H. Davy isolated the impure metal in 1808, he called it "magnium." At that time, the terms "magnesium" and "manganesium" were applied synonymously to the element (manganese) derived from the mineral pyrolusite (manganese dioxide). To avoid confusion, the term "magnesium" was soon afterwards restricted to the element derived from magnesia alba; and "manganese" to the element derived from pyrolusite.¹

Occurrence of magnesium.—Magnesium occurs in nature combined, not free, as magnesium carbonate in *magnesite*, MgCO_3 ; double carbonate of calcium and magnesium in *dolomite*, $\text{MgCO}_3 \cdot \text{CaCO}_3$; magnesium sulphate in *Epsom salts*, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$; and *kieserite*, $\text{MgSO}_4 \cdot \text{H}_2\text{O}$; magnesium chloride in *carnallite* and *kainite*; magnesium silicate in *olivine*, Mg_2SiO_4 ; *enstatite*, MgSiO_3 , etc. Magnesium is also common in many other minerals: e.g. asbestos, steatite, spinel, meerschaum, serpentine, talc, etc.

Atomic weight.—The combining weight of magnesium (oxygen = 16) has been determined by the analysis of the sulphate, oxalate, chloride, etc., and some of the best results lie between 24.26 and 24.39. The best representative value is taken to be 24.32. This agrees with the value for the atomic weight estimated by Dulong and Petit's method of approximation, specific heat, 0.2234; and the isomorphism of some magnesium salts with iron, zinc, manganese, etc., salts. Beryllium has an atomic weight approaching 9.1.

Preparation of the metals.—Magnesium and beryllium are closely related metals, both can be prepared by the electrolysis of the chlorides; or of a mixture of the respective chlorides with potassium chloride. In the case of magnesium, fused carnallite can be used. The metals can also be made by heating the chlorides with sodium: $\text{MgCl}_2 + 2\text{Na} = 2\text{NaCl} + \text{Mg}$. Both processes—electrolysis and sodium reduction—are used for magnesium on a large scale. Beryllium is of little commercial importance.

Laboratory preparation of magnesium.—An artificial carnallite is first made by evaporating a solution of 20 grams crystalline magnesium chloride, 7.5 grams of potassium chloride, and 3 grams of ammonium chloride to dryness in a platinum dish on a water bath, and fuse the residue over a blast gas burner (Fig. 143). Pour the molten liquid into the bowl of a clay pipe (Fig. 143) heated over a gas burner. Put a knitting-needle down the stem of the pipe, and sink a carbon rod into the bowl as shown in the diagram. Connect the steel needle and the carbon rod with a source of electricity so that the former can act

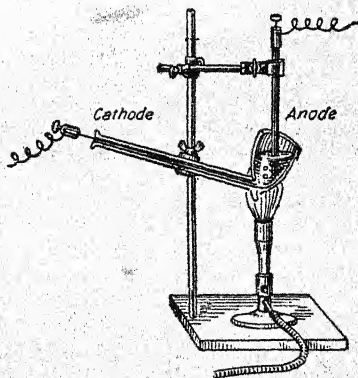


FIG. 143.—Electrolysis of Fused Carnallite.

¹ It is not very clear whether the term "magnesia" is a corruption of the word *Mangana* in the East Indies, or whether "manganese" is a corrupted form of the word *Magnesia*, a locality in Asia Minor.

as cathode, the latter as anode. Cover the molten mass with a layer of charcoal dust. Pass a current of electricity (about 10 volts and 8 amps.) through the molten mass for about half an hour, and let everything cool. When the mass is broken up in a mortar, the small globules of magnesium metal, scattered here and there throughout the mass, can be picked out, and washed clean by means of alcohol. The object of the charcoal powder is to protect the metal from oxidation should any globules rise to the surface.

Properties of magnesium and beryllium.—Both metals have a silvery-white lustre, and low specific gravity. When ignited, magnesium burns in air, giving a brilliant white light of great actinic power. Both metals are slowly oxidized by moist air. Magnesium very slowly decomposes boiling water. When heated in a current of steam, magnesium takes fire, and continues burning: $\text{Mg} + 2\text{H}_2\text{O} = \text{Mg}(\text{OH})_2 + \text{H}_2$. Magnesium melts at 651° , beryllium at about 1280° . Magnesium boils at about 1120° . Both metals readily dissolve in dilute hydrochloric and sulphuric acids; magnesium is also rapidly dissolved by nitric acid, and it is perhaps the only metal which gives a copious yield of hydrogen when treated with nitric acid. Beryllium is not readily attacked by cold nitric acid, but is readily attacked by the hot acid. Beryllium is dissolved by cold concentrated (say, 50 per cent.) solutions of the alkaline hydroxides and hydrogen gas is evolved: $\text{Be} + 2\text{KOH} = \text{Be}(\text{OK})_2 + \text{H}_2$; a more dilute (say, 10 per cent.) solution attacks the metal only when heated. Alkali hydroxides thus behave towards the metal like a weak acid. Magnesium is not appreciably attacked under the same conditions. Beryllium seems to bear the same relation to alkaline earth metals that lithium bears to the alkali metals. Magnesium reacts with aqueous solutions of ammonium salts, forming a double salt with the evolution of hydrogen. Magnesium when heated with nitrogen forms magnesium nitride, Mg_3N_2 . Magnesium is used in flashlight preparations for photography, and also in making fireworks.

Magnesium and beryllium oxides.—The oxides can be prepared by igniting the metals in air, as well as by calcining the nitrates, carbonates, and sulphates in the presence of water vapour. Magnesium oxide is made from magnesium chloride of the Stassfurt deposits by converting the latter into the carbonate and calcining the resulting carbonate. The "magnesia usta" of commerce is made by the prolonged calcination of the carbonate at a low temperature. The so-called heavy magnesia, or *magnesia ponderosa*, of commerce is made by calcining heavy magnesium carbonate, and the light magnesia, or *magnesia levis*, by calcining light magnesium carbonate. The two powders have a different specific gravity.

Magnesia is slightly soluble in water, to which it gives a slight alkaline reaction: 100 c.c. of water dissolve about 0.001 gram of magnesia. The higher the temperature of calcination, the less the solubility of the oxide, due, it is said, to the polymerization of the magnesia. This, however, has not been proved; the alleged lower solubility may simply mean that the rate of dissolution is reduced. The oxide is not completely converted into the hydroxide by the action of water. The hydroxide is precipitated from solutions of magnesium salts by the addition of alkali hydroxides. Magnesium hydroxide, $\text{Mg}(\text{OH})_2$, is soluble in ammonium salts, and hence the precipitation with ammonia is incomplete: $\text{Mg}(\text{OH})_2 + 2\text{NH}_4\text{Cl} \rightleftharpoons \text{MgCl}_2 + 2\text{NH}_4\text{OH}$. If enough ammonium salt be present in the solution, magnesium hydroxide will not

be precipitated at all; at the same time, a soluble salt—probably $(\text{NH}_4)_2\text{MgCl}_4$, that is, $\text{MgCl}_2 \cdot 2\text{NH}_4\text{Cl}$ —is formed in the solution.

Magnesium oxide fuses at about 2500° , and beryllia at about 2400° . On account of the refractory qualities of magnesia, it is used for pencils as an alternative to lime, in the so-called "lime-light"; and also for making refractory bricks ("magnesia" bricks made from calcined magnesite, and "dolomite bricks" from calcined dolomite), crucibles, cupels, furnace linings, etc. A paste made with water and magnesia calcined at a low temperature behaves similarly to lime in mortar. It gradually re-hydrates, absorbs carbon dioxide from the air, and sets in about twelve hours to a hard mass.

§ 2. Zinc, Cadmium, and Mercury—Occurrence and Preparation.

History of zinc.—Brass, an alloy of zinc and copper, was known to the ancients, and several references to "brass" occur in the sacred writings. R. Jagnaux says that bracelets made of zinc have been found in the ruins of Cameros which was destroyed about 500 B.C. Such knowledge of zinc as was possessed by the ancients appears to have been lost for a time. The word "zinken" appears in the writings of Basil Valentine, but he did not refer to it as a metal, and he seems to have confused it with several other substances. Paracelsus also used the word in the 16th century. At that time the term *zinc* appears to have been employed locally by the miners in the Carnia for an earthy ore. Agricola, in 1596, said that the smelters of Goslar extracted a metal from an ore at Rammelsberg. Near the beginning of the 17th century, zinc was brought from China and the East Indies under the name "tutanego." Indian tin, calamine, or speauter. Libavius (1597) said that the Indian tin and the Goslar metal were the same. About this time, however, there was much confusion as to the meaning of zinc. In 1697, Löhneyes appears to have been the first to definitely apply the term "zink" to the metal now known as zinc. In 1695 Mr. Homberg, and in 1721 J. F. Henkel, discovered that zinc could be obtained from calamine, and a works for the manufacture of zinc was erected at Bristol about 1740 by J. Champion. Champion's process was patented 1739; in this process the ore was distilled in large crucibles arranged with a pipe extending downwards through the bottom of the crucible. It is called the English process of "distillation per descensum." The process is obsolete.

Occurrence of zinc.—Metallic zinc has been reported in the basaltic rocks of Victoria (Australia); but it usually occurs combined: as carbonate, *zinc spar*, *calamine*, ZnCO_3 ; sulphide, *zinc blende* or *black jack*, ZnS ; oxide, *zincite* or *red zinc ore*, ZnO ; silicate, *willemite*, $2\text{ZnO} \cdot \text{SiO}_2$; *franklinite*, $(\text{ZnFe})\text{O} \cdot \text{Fe}_2\text{O}_3$; *zinc spinel* or *gahnite*, $\text{ZnO} \cdot \text{Al}_2\text{O}_3$.

History of cadmium.—The term *καδμεία* (*cadmeia*) was applied by Discorides, and by Pliny, to a zinciferous earth (*calamine*)—found on the shores of the Black Sea—which when melted with copper furnished brass—aurichalcum. Pliny also applied the term "*cadmia*" to the tutty (impure zinc oxide) found in the flues of brass-founders' furnaces. In 1817 F. Stromeyer discovered a yellow oxide free from iron in a sample of zinc carbonate used at the smelting works at Salzgitter. This could only be due to the presence of a new metal which he called "*cadmium*," from

cadmia fornacum, because the metal was found in the "flowers of zinc," that is, the flue dust of the zinc furnace.

Occurrence of cadmium.—This element does not occur free. It is commonly found accompanying zinc in calamine, and zinc blende. Very few zinc ores contain more than 0.5 per cent. of cadmium. The rare mineral *greenockite*, cadmium sulphide, CdS , is of no commercial importance.

Preparation of zinc.—Zinc and cadmium usually occur together. The ores employed are the carbonate, oxide, and sulphide. The process involves two operations: (1) *Roasting*.—The object is to convert the sulphide into the oxide by calcination. The oxidation of the carbonate presents no difficulty: $\text{ZnCO}_3 = \text{ZnO} + \text{CO}_2$. In the case of the sulphide, the sulphur is oxidized by calcination in air, usually in furnaces with mechanical stirring devices, $2\text{ZnS} + 3\text{O}_2 = 2\text{ZnO} + 2\text{SO}_2$. (2) *Reduction*.—The crushed oxide is mixed with crushed coke and heated to bright redness in fireclay retorts. The oxide is reduced with the formation of carbon monoxide: $\text{ZnO} + \text{C} = \text{Zn} + \text{CO}$. The common oxides are reduced by a reaction of the type $\text{MO} + \text{CO} \rightleftharpoons \text{M} + \text{CO}_2$, which is reversible; but the concentration of the carbon dioxide must be very great before the right-to-left reaction becomes appreciable. With the reaction between zinc vapour and carbon dioxide, the back reaction is marked; and when carbon dioxide is passed over hot zinc, the gas is reduced to carbon monoxide, and zinc oxide is formed. Hence, to complete the reduction of zinc oxide in the retort, it is necessary to continuously destroy the carbon dioxide by using an excess of carbon: $\text{CO}_2 + \text{C} = 2\text{CO}$. Zinc oxide is reduced by carbon at a temperature higher than the volatilization temperature of the zinc. The metal distils over at about 1030° , and collects in fireclay or iron receivers. At first a fine bluish-grey powder—"zinc dust"—collects in the receiver. This is a mixture of zinc oxide and powdered metal. When the receiver is warm, the metal condenses to a liquid, which is drawn off at intervals and cast into plates or bars. If zinc sulphate had been produced during the roasting of the sulphide, $\text{ZnS} + 2\text{O}_2 = \text{ZnSO}_4$, it would be reduced by the carbon back to the sulphide, and thus reduce the yield of metal by the ore. The zinc so obtained—called *spelter*—contains carbon, iron, lead, arsenic, and cadmium as impurities. It may be refined by careful distillation. The fireclay retorts have different shapes, and different types of condensers are used in different smelting works. The prevailing styles are the Silesian, Belgian, and Rhenish. The so-called Belgian retort is an oval or cylindrical tube—6" to 10" diameter, and about 3' or 4' long. The retort is fitted with a conical tube—about 16" long—as condenser, Fig. 144. The wide end fits the retort and the narrow end is fitted with a sheet iron nozzle to catch any

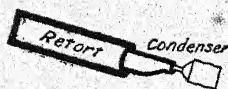


FIG. 144.—Belgian Zinc Retort.

zinc dust. The retorts are built into the furnace in rows and tiers. Some of the furnaces running to-day have "banks" with hundreds of retorts. There have been no radical changes in the metallurgy of zinc since the first Belgian furnaces were built in the early part of last century. Electric furnaces for the smelting of zinc ores have made no pronounced progress, mainly due to the fact that a large unmanageable proportion of zinc dust is produced.

Preparation of cadmium.—The first product of the distillation of zinc ores contains most of the cadmium partly as metal, partly as oxide. The zinc dust is reduced in clay or iron retorts and redistilled. The product, called "crude cadmium," contains some zinc from which it can be separated by repeated distillation at a low temperature or by electrolysis. A solution of cadmium chloride or sulphate—containing up to 30 per cent. of cadmium—is electrolyzed with crude cadmium as the anode, and pure cadmium as cathode. The cadmium dissolves from the former and is redeposited on the latter. There are not enough cadmium ores to allow cadmium to be extracted for its own sake, and hence the extraction of cadmium is often a "side line" in a zinc smelting works.

History of mercury.—Mercury was mentioned 300 B.C. by Theophrastus as *χυτός ἀργυρος* (chutos argyros), quicksilver or liquid silver; and he states that it can be made by rubbing vinegar with cinnabar in a copper vessel. Dioscorides called it *ὑδωρ ἀργυρος* (hydor argyros), liquid silver; hence the Latin *hydrargyrum*, and the present-day symbol Hg. The metal had a certain fascination for the alchemists, and for a time they believed that it, or something similar, was a constituent of all metals. "Nimble volatile mercury" was named after the mythological Mercury, the messenger of the gods, and accordingly the ancient chemists symbolized the metal by the caduceus or herald's wand ☿, also used for the planet Mercury.

Occurrence of mercury.—Free mercury in small quantities occurs disseminated in the ores of mercury, which in turn occur in relatively few places in payable quantities. The mercury deposits are usually found along lines of profound volcanic disturbances. *Cinnabar*, HgS, is the chief ore of mercury, and it is mined in Almaden (Spain), Idria (Carniola), Bavarian Palatinate, Peru, California, Japan, China, etc. The great quicksilver mine at Almaden is said to have been worked at least as far back as 415 B.C. Pliny (c. A.D. 77) reported that 10,000 lbs. of cinnabar were brought to Rome per annum from this locality. The American deposits range from Alaska along the Pacific slope down to Peru. At one time the output from America rivalled that from Almaden, but, unlike the latter deposits, the American, and indeed most others, decrease in value with increasing depth.

Preparation of mercury.—Mercury is obtained almost exclusively from cinnabar, HgS. The ore is first crushed dry and then sorted by hand or screen into two grades. The one is a rich ore, the other contains much gangue. The ore is roasted to oxidize the sulphur, and the metal is liberated: $\text{HgS} + \text{O}_2 = \text{SO}_2 + \text{Hg}$; or else the ore is distilled with iron or lime in closed retorts whereby calcium sulphide and mercury are formed: $2\text{HgS} + 2\text{CaO} = 2\text{CaS} + 2\text{Hg} + \text{O}_2$; or $\text{HgS} + \text{Fe} = \text{FeS} + \text{Hg}$. The reactions here take place at a temperature higher than the boiling point of mercury, so that the metal separates in the vaporous state. The greatest difficulty in the extraction of mercury lies in its extreme volatility. The workmen, too, have to be protected from the toxic vapour. The air oxidation process is generally used with rich ores, but different condensing arrangements are employed in different works. The mercury, for example, may be condensed in large chambers as at Idria; or in a series of pear-shaped vessels—aludels—connected in rows nearly 50 feet long as at Almaden. About six trains of aludels are con-

nected with one roasting furnace. With poor ores, the roasting is conducted in reverberatory furnaces. Vertical furnaces are also employed; these are fed at the top alternately with ore and fuel. The sulphurous and mercurial vapours pass from the furnace through condenser tubes cooled by a stream of water on the outside. Crude mercury is sent into commerce in iron bottles holding about 75 lbs. of liquid metal. The mercury may be cleaned by filtration through chamois leather; and purified by distillation from iron retorts. In the laboratory, mercury is often purified by running a fine spray of mercury down a long column of dilute nitric acid (specific gravity 1.1), or a solution of mercurous nitrate, followed by distillation *in vacuo*. Purified mercury does not "wet" glass, but runs over glass in clean rounded globules; if, however, the mercury be contaminated with foreign metals, a film of oxide is formed and the globules which run on the glass produce a "tail" on the surface of the glass.

§ 3. The Properties of Zinc, Cadmium, and Mercury.

Cadmium and zinc are white metals. If zinc be melted in a crucible and poured on to a non-conducting surface of, say, asbestos, and the liquid portion then run off from that which has solidified, a crust of hexagonal pyramidal crystals will be obtained. Obviously the external geometrical form of a crystal can be developed only when the crystal is developed in a free liquid. When cooling a liquid solidifies *en masse*, the growing crystals interfere with one another and, when a section is cut through the mass, the solid appears as an aggregate of irregular polygons. Most metals are aggregates of minute crystals grown under these conditions. If the polished surface of a section through a piece of tin be etched with hydrochloric acid; zinc, with very diluted hydrochloric acid; or a piece of lead with diluted nitric acid, the polygonal areas can be seen with the naked eye; with iron, the surface must be magnified because the crystals are so small. In general, for a given metal, the average size of the crystals depends upon the rate of cooling. The slower the cooling the larger the crystals. Thus, the centre of a large mass of cast metal will appear to have a more coarsely granular texture than near the boundary walls. If a solution of one metal in another be cooled, one constituent may crystallize out and leave the eutectic mixture to solidify last. On the polished or etched surface of such an alloy, the crystals appear to be embedded in the eutectic matrix; illustrations appear in Figs. 59, 64, etc. Zinc is brittle at ordinary temperatures, but ductile at 100°–150°, and it can then be drawn into wire, and rolled into sheets. Cadmium is ductile enough at ordinary temperatures to be rolled into sheets and drawn into wire. Cadmium wire when heated for a long time to 80° disintegrates. This is supposed to be due to the formation of an allotropic modification which has a transition temperature of 65°. If ordinary cadmium be heated for, say, 24 hours at this temperature, the specific gravity diminishes. Zinc, similarly treated at 100° changes in specific gravity from 7.129 to 7.102. Hence it is also inferred that zinc exists in at least two allotropic forms. Zinc is brittle at 200°, and it can then be readily powdered. Cadmium melts at 320.2°, zinc at 419.4°. Mercury is a silvery-white metal with a slight blue tinge; it is liquid at ordinary temperatures; caesium melting

at about 26° may be liquid on a hot summer's day. In thin films, mercury is violet by transmitted light. It solidifies into a malleable solid at -38.87° . Mercury does not tarnish in air. Both cadmium and zinc are slowly oxidized by moist air, and in water containing air in solution, the metals are oxidized with the formation of basic carbonates. Both metals are attacked by dilute hydrochloric and sulphuric acids giving hydrogen, and a zinc or cadmium salt. The reaction with sulphuric acid is influenced by the temperature and concentration of the acid; thus, with nearly pure zinc and dilute acid, the products are hydrogen and zinc sulphate; as the temperature or concentration of the acid rises, the proportion of hydrogen decreases, and hydrogen sulphide and sulphur dioxide appear, until, with very concentrated acid, sulphur dioxide is the main gaseous product. With commercial zinc of a low degree of purity, sulphur dioxide and hydrogen sulphide can generally be detected, even when the acid has a dilution of 1:100. Hence the reaction between zinc and sulphuric acid is more complex than is represented by the equation $\text{Zn} + \text{H}_2\text{SO}_4\text{aq.} = \text{ZnSO}_4\text{aq.} + \text{H}_2$. Nitric acid and zinc give oxides of nitrogen. Cadmium is insoluble in alkaline hydroxides, but zinc dissolves, giving off hydrogen (*q.v.*). Mercury is not attacked by hydrochloric acid; concentrated sulphuric acid acts very slowly in the cold; but when heated, mercuric sulphate, HgSO_4 , sulphur dioxide, and some mercurous sulphide are formed. Concentrated nitric acid rapidly attacks mercury, forming mercuric nitrate and oxides of nitrogen. Dilute nitric acid acts slowly, giving mercurous nitrate. Alkali hydroxides have no appreciable action on mercury.

Amalgams.—Mercury is a good solvent for some of the metals. The solutions are called "amalgams." Mercury amalgamates, or, in the language of the alchemists, mercury devours or licks up gold, silver, cadmium, tin, bismuth, lead, zinc, etc., at ordinary temperature, and others are dissolved at higher temperatures. Finely divided copper readily amalgamates, but in mass the action is very slow. Arsenic, antimony, and platinum can be amalgamated with difficulty, while cobalt, nickel, and iron do not amalgamate directly. The phenomena attending the solution of the metals in mercury appear to be closely analogous with the solution of different substances in water. A considerable amount of heat is often developed, as is the case when sodium or potassium metals are dissolved in mercury. The freezing points of solutions of potassium in mercury are indicated by the curve, Fig. 145.

Compare Figs. 59, 104, etc. The breaks in the curve correspond with the formation of the following compounds: KHg (melts 178°); KHg_2 (270°); K_2Hg_9 (204°); K_2Hg_3 (173°); KHg_3 (70°). Two well-defined eutectics

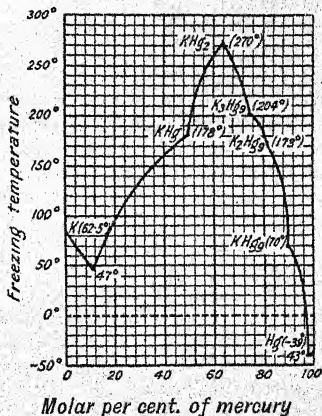


Fig. 145.—Freezing Curves of Potassium-Mercury Amalgams.

occur at 47° and -43° . The numbers in brackets represent melting points. With sodium, the following compounds are indicated on the freezing-point curve: NaHg_4 (159°); NaHg_2 (360°); NaHg (129°); Na_3Hg_2 (123°); Na_2Hg_2 (62°); Na_3Hg (-34°). When the potassium amalgams are heated to 440° , crystalline KHg_2 remains behind, with sodium, Na_3Hg . These latter compounds spontaneously inflame when exposed to the air. Gold and silver dissolve readily in mercury, and this fact is utilized in the extraction of gold (*q.v.*), when the mercury acts as a "collector and saver" of minute particles of gold which would otherwise be lost. Hence, it has been said that "mercury is practically indispensable in the gold-mining industry." The cyanide process has, however, weakened the force of that quotation. Tin amalgam is soft, and is used in making mirrors. Amalgams of gold, copper, and zinc, are used in dentistry for stopping teeth. Zinc amalgam is but slowly attacked by sulphuric acid, and amalgamated zinc is used in making batteries, because the zinc is not then wasted by dissolution in the acid; the amalgamated zinc dissolves only when the circuit is closed. To amalgamate zinc plates, first clean the plate free from grease by washing it with dilute soda-lye; then wash in dilute sulphuric acid; and finally rub mercury well over the surface until the zinc plate appears to be covered with a bright silvery film.

The atomic weight of zinc and cadmium.—The combining weight of zinc has been determined by the analysis of the halogen compounds, the carbonate, and the sulphate, and the synthesis of the oxide. Some of the best results (oxygen = 16) vary between 65.24 and 65.99; and the best representative value is taken to be 65.37, which agrees with the atomic weight estimated from the isomorphism of zinc salts with some salts of magnesium, manganese, etc.; the vapour density of volatile compounds of zinc; and Dulong and Petit's method of approximation, specific heat of zinc, 0.0935. The atomic weight of cadmium determined by similar methods is 112.4. Zinc and cadmium are usually considered bivalent, although organic chemists recognize the probable existence of monovalent forms.

The atomic weight of mercury.—Some of the best determinations of the combining weight of mercury, by the analysis or synthesis of mercury oxide, sulphide, chloride, bromide, and cyanide give numbers ranging from 199.83 to 200.23. The best representative value (oxygen = 16) is taken to be 200; and this number also corresponds with the atomic weight estimated from the vapour density of volatile compounds of mercury; and the isomorphism of some mercury salts with salts of lead, copper, etc.

Uses.—Zinc is used in making the anode plates of batteries and in making certain utensils. It is also a prevailing constituent in certain alloys: brasses, German silver, bronze, etc. (*q.v.*). Galvanized iron is iron covered with a protective coating of zinc to prevent rusting. In one process of galvanizing, the iron is first cleaned with acid or sand blast, and subsequently dipped in molten zinc; in another process of galvanizing, the zinc is deposited electrolytically, similar to electroplating, described on p. 357. Mercury is used in making certain medicinal preparations—blue pills, mercurial ointments, corrosive sublimate, calomel, etc.; in making the pigment vermilion; in the manufacture of mercury fulminate—the active agent used in igniting detonating mixtures and explosives; in making amalgams; and in the manufacture of scientific instruments—

thermometers, barometers, etc. Cadmium is used in making certain fusible alloys (*q.v.*).

§ 4. The Oxides of Zinc, Cadmium, and Mercury.

Zinc monoxide, ZnO ; and cadmium monoxide, CdO .—The monoxides of zinc and cadmium are formed when the metals are burnt in air: $2\text{Zn} + \text{O}_2 = 2\text{ZnO}$. Zinc oxide, under the commercial name "zinc white," is manufactured by heating zinc in air and passing the fumes into condensing chambers where the oxide collects. Zinc oxide is used in place of white lead as a white pigment where the blackening of white lead is objectionable; nor is zinc oxide liable to produce toxic effects if carelessly used, as is the case with lead paints which then produce "painter's colic." A large quantity of zinc oxide is used in the manufacture of rubber goods—tyres for motor cars, etc. Zinc oxide appears yellow when hot, white when cold. The mineral *zinkenite* is a form of zinc oxide which occurs in hexagonal prisms. Cadmium oxide, CdO , has a rich brown colour. Zinc oxide forms hexagonal crystals if heated to a very high temperature; it does not melt even in the oxyhydrogen blowpipe, but, like lime and magnesia, the oxide is vividly incandescent under these conditions, it afterwards appears phosphorescent in the dark.

Zinc hydroxide, $\text{Zn}(\text{OH})_2$; cadmium hydroxide, $\text{Cd}(\text{OH})_2$.—These hydroxides cannot be produced by the action of water on the oxides. As in the case of magnesium salts, the hydroxide is precipitated by ammonia, but it is soluble in an excess, particularly in the presence of ammonium salts. 100 c.c. of water dissolve but 0.0004 gram of zinc oxide. The hydroxides are precipitated when an alkaline hydroxide is added to a solution of a zinc or a cadmium salt. Zinc hydroxide, unlike cadmium hydroxide, is soluble in an excess of the alkaline hydroxide, forming salts of the type $\text{Zn}(\text{OK})_2$, called zincates. The same salt is formed when zinc metal is dissolved in potassium hydroxide. The zincates are to be regarded as salts of the amphoteric zinc hydroxide, $\text{Zn}(\text{OH})_2$, which behaves as a dibasic acid, zincic acid, H_2ZnO_2 , and as a basic hydroxide. Both hydroxides are soluble in aq. ammonia, and ammonium chloride, not because of the acidic character of the hydroxide, but because of the formation of soluble complex compounds with ammonia—*e.g.* the solution of cadmium hydroxide in aq. ammonium chloride furnishes a mixture of crystals of $\text{CdCl}_2 \cdot 2\text{NH}_3$; $\text{CdCl}_2 \cdot \text{NH}_4\text{Cl}$; and $\text{CdCl}_2 \cdot 4\text{NH}_4\text{Cl}$. Both oxides are basic and yield salts on treatment with acids.

Zinc and cadmium peroxides.—Both zinc and cadmium form peroxides: ZnO_2 and CdO_2 , when the oxides are moistened with hydrogen peroxide. Better yields are obtained by electrolyzing solutions of the respective chlorides in vessels fitted with a porous diaphragm and with hydrogen dioxide added to the cathode compartment. They are probably superoxides, so that they are represented by the constitutional formulæ:

$\text{Zn} \begin{smallmatrix} \text{O} \\ \diagup \quad \diagdown \\ \text{O} \end{smallmatrix}$ and $\text{Cd} \begin{smallmatrix} \text{O} \\ \diagup \quad \diagdown \\ \text{O} \end{smallmatrix}$. These peroxides are easily decomposed by acids with the evolution of oxygen. The peroxides of zinc and cadmium, and also of magnesium are less stable than the peroxides of the alkaline earths.

Cadmous oxide, Cd_2O .—When cadmium chloride, CdCl_2 , is heated with metallic cadmium, a lower chloride, cadmous chloride, CdCl , is

formed. This when treated with water gives cadmous hydroxide, CdOH , and this, in turn, when gently warmed, gives cadmous oxide, Cd_2O . Zinc does not form an oxide lower than the monoxide, ZnO ; hence cadmium, in this respect, is more closely related with mercury, which also forms two oxides, mercuric oxide, HgO , and mercurous oxide, Hg_2O .

Mercurous oxide, Hg_2O .—This oxide is formed as an unstable dark brown powder when calomel, HgCl , is digested with an alkaline hydroxide: $\text{HgCl} + \text{KOH} = \text{KCl} + \text{HgOH}$; $2\text{HgOH} = \text{Hg}_2\text{O} + \text{H}_2\text{O}$; and as a black powder when an alkaline hydroxide is added to soluble mercurous salts. No corresponding hydroxide is known. Instead of being oxidized in air, mercurous oxide is decomposed when warmed, or when exposed to air in the light, forming mercuric oxide and mercury: $2\text{Hg}_2\text{O} = 2\text{HgO} + 2\text{Hg}$. Mercurous oxide is feebly basic, and it exhibits no acidic qualities like zinc oxide.

Mercuric oxide, HgO .—At first sight, this oxide appears to exist in two distinct modifications—red and yellow. If a mercuric salt be treated with an excess of alkaline hydroxide in the cold, mercuric hydroxide is probably formed, but this immediately decomposes into yellow mercuric oxide. If the precipitation be made from hot solutions, an orange precipitate is obtained. The difference in colour is probably due to the state of subdivision of the precipitate. The larger the granules, the redder the tint. If mercuric oxide be prepared by the ignition of the nitrate, or by calcining the metal in air, nearly at its boiling point, the oxide is red and distinctly crystalline, the yellow oxide becomes red when heated to about 400° . When heated, the red oxide darkens in colour and finally appears almost black; the red colour returns on cooling. If heated above the temperature at which the oxide appears to blacken, it decomposes into mercury and oxygen. The yellow and red oxides are slightly soluble in water. At 25° a litre of water dissolves 0.0513 gram of the red and 0.0518 gram of the yellow oxide; the solution has an alkaline reaction with basic, but no acidic, qualities.

On account of the ease with which mercuric oxide parts with its oxygen, it is an active oxidizing agent. The yellow oxide is more active than the red, probably owing to its finer state of subdivision. Hence, the yellow oxide was used in preference to the red in preparing chlorine monoxide and hypochlorous acid.

§ 5. The Magnesium-Zinc Family of Elements.

These elements form a family related in many ways with one another, and with the metals of the alkaline earths. Beryllium and magnesium form a kind of subgroup; zinc, cadmium, and mercury form another subgroup. The metals beryllium and magnesium, but more particularly the former, are "bridge elements" in that they appear to link the alkaline earths with zinc, cadmium and mercury. The scheme indicated in the margin is sometimes used to illustrate the idea. The vapours of all the elements appear to be composed of monatomic molecules. The chemical relations have been discussed in what precedes; the physical properties of the metals are summarised in the following table:—

Be
|
Mg
/ \
Ca Zn
| |
Sr Cd
| |
Ba Hg

TABLE XXII.—PHYSICAL PROPERTIES OF THE MAGNESIUM-ZINC METALS.

—	Beryllium.	Magnesium.	Zinc.	Cadmium.	Mercury.
Atomic weight .	9.1	24.32	65.37	112.40	200.0
Specific gravity .	1.64	1.75	6.9-7.2	8.6	13.6
Atomic volume .	5.5	13.8	9.34	13.0	15.4
Melting point .	1280°	651°	419.4°	320.9°	-38.7°
Boiling point .	—	1120°	916°	780°	357.3°

The metals are not oxidized so readily as the alkaline earths. The affinity of the metals for oxygen decreases with increasing atomic weight. Beryllium does not bear so close a relationship to magnesium, zinc, and cadmium, and, while mercury has a great many similarities, it has many important differences, thus: (1) The salts are all volatile; (2) it does not readily combine with oxygen; (3) its hydroxide is difficult to make; (4) the black sulphide is virtually insoluble in nitric acid; and (5) it forms two chlorides one of which resembles silver chloride. As a matter of fact, the properties of the mercuric salts can scarcely be said to fraternize very closely with the salts of any other metal.

The carbonates of these elements break up when heated into carbon dioxide, and a residual oxide, and the oxides are but sparingly soluble in water. Magnesium oxide is white; zinc oxide is white when cold, yellow when hot; cadmium oxide is brown when cold; and mercuric oxide is red. The oxides and hydroxides are soluble in solutions of ammonium salts. Cadmium, like mercury, forms a lower oxide and chloride. The basic character of the hydroxides decreases with increasing atomic weight. The hydroxides are not made by direct union of the oxide with water, and the water is easily expelled from the hydroxides by heat. This is not the case with the hydroxides of the alkaline earths. The sulphates are soluble, and zinc and magnesium sulphates are isomorphous; the sulphates are all less stable than those of the alkaline earths, and their stability decreases with increasing atomic weight. The sulphates all combine with potassium sulphate—*e.g.* $\text{ZnSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$. The chlorides are all volatile deliquescent solids; zinc and magnesium chlorides are readily hydrolyzed in aqueous solution, and form basic salts when the solutions are evaporated to dryness. This is not the case with cadmium and mercuric chlorides. The two latter readily combine with ammonia to form complex salts, and cadmium like mercury forms a lower oxide and chloride; the iodides of these two elements are isomorphous. The halides become less stable on passing from beryllium to mercury. Magnesium and zinc chlorides are hydrolyzed in aqueous solution and form basic salts on evaporation to dryness. This is not the case with solution of cadmium and mercuric chlorides. The sulphides increase in stability and are less soluble as the molecular weight increases. The sulphides of the first two elements are not formed in the presence of water; zinc sulphide is stable in aqueous and slightly acidic solutions; cadmium sulphide is dissolved by concentrated and not by dilute acids; whereas mercuric sulphide is scarcely attacked, even by boiling nitric acid.

Questions.

1. State any facts known to you which tend to show that some alloys are compounds and not mere mixtures of the constituent metals.—*London Univ.*

2. One litre of mercury vapour at the standard temperature and pressure weighs 8.923 grams. On heating 118.3938 grams of mercuric oxide, Erdmann and Marchand obtained 109.6308 grams of mercury. On the assumption that mercuric oxide is formed by the union of one atom of mercury with one atom of oxygen, what light do these facts throw on the atomic and molecular weights of mercury?—*Science and Art Dept.*

3. How is potassium chloride converted into (1) caustic potash, (2) potassium chlorate? Starting from carnallite, the double chloride of magnesium and potassium, show briefly how the metals magnesium and potassium can be obtained.—*Owens Coll.*

4. How can mercuric and mercurous chloride be obtained from mercuric sulphate? What is the action of mercuric chloride solution with solutions of (a) potassium iodide, (b) stannous chloride, (c) ammonia, (d) sodium hydroxide?—*London Univ.*

5. How would you prepare pure calcium carbonate from an impure magnesium limestone?—*Cleveland Univ. School, Ohio.*

6. The adjoining curve, Fig. 146, represents the melting points of the series of alloys formed by two metals. What do you learn from the curve, and what would be the appearance under the microscope of prepared specimens of the compositions marked A, B, C?—*Sydney Univ.*

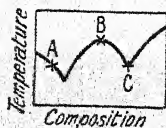


FIG. 146;

7. Mercury forms two classes of salts—mercurous and mercuric salts. Illustrate by examples how these salts differ in composition from each other. How would you discover whether a given mercury salt was a mercurous or a mercuric salt?—*London Univ.*

8. A crystalline compound afforded the following numbers on analysis:—

Magnesium	Potassium	Sulphur	Oxygen	Water
5.95	19.44	15.99	31.82	26.86 (Total 100.00).

Calculate its formula, given $Mg = 24$, $K = 39$, $S = 32$, $O = 16$, $H = 1$.—*London Univ.*

9. Six grams of anhydrous magnesium sulphate are dissolved in 100 grams of water; the observed depression of the freezing point is 0.958° . Taking the value of the freezing constant as 18.5, what formula most probably represents the state of the dissolved salt?

10. Supposing the composition of lime to be still unknown, show that a comparison of the properties of lime with those of the oxides formed on burning magnesium and zinc in air would make it appear probable that lime also is a metallic oxide.—*London Univ.*

11. Give an account of the metal mercury and of its behaviour towards the common acids. Give the formulae and names (systematic and trivial) of its chief compounds.—*London Univ.*

12. A zinc blende contains 72 per cent. of zinc sulphide, ZnS , and 0.5 per cent. of cadmium sulphide, CdS ; the rest consists of refractory non-reducible earthy matters. How much coke containing 92 per cent. carbon is theoretically necessary for reducing the roasted blende, and how much zinc and cadmium can be theoretically obtained from 2 tons of the original zinc blende?—*G. Bornemann, Stochiometrie, Leipzig, 1909.*

13. What are the chief ores found in Europe from which zinc is extracted? Describe the process used in England and Belgium for the extraction of zinc and for its purification. What other less common elements are found associated with zinc, and how are they separated? What are the common impurities in commercial zinc?—*Science and Art Dept.*

CHAPTER XIX

THE ALKALI METALS

§ 1. Potassium and Sodium Carbonates.

Potash is both wealth and weapon. Our position as sole producer gives us the power of causing grave injury to the agriculture of our enemies by stopping the export of salts so indispensable to them for manurial purposes. If raw materials are refused us, we shall revenge ourselves on the enemy's agriculture with this war cry: "At the enemy with a kilo of potash!"—*VOSSISCHE ZEITUNG* (Nov. 23, 1917).

POTASH is an indispensable agent for manurial purposes in agriculture, and during the Great War, when supplies from the Stassfurt deposits were cut off, a serious potash famine was threatened. The chief sources of potash are: (1) Natural potash beds at Stassfurt (Germany), Alsace (France), Galicia (Austria), Catalonia (Spain), Punjab (India), Atacania (Peru), Searle's lake (Nebraska), etc. (2) Sea-water, brines, and many lake deposits have potash salts associated with sodium salts which can be separated by fractional crystallization. (3) The ashes of vegetable matter—*e.g.* wood ashes, beetroot residues, seaweeds, sunflower stalks, hedge trimmings, banana skins, etc. (*vide infra*). (4) The water used in washing grease from wool (*vide infra*). (5) Products resulting from the decay of organic matter—*e.g.* Indian and South African nitres. (6) Blast furnace and cement-kiln dust. (7) The insoluble potash minerals—*e.g.* felspar, alunite, leucite, etc.

The ash of wood, not coal, contains about 5 to 15 per cent. of potassium carbonate. In special districts it may be profitable to burn wood in pits and extract the ashes with water in wooden tubs. The clear liquid is evaporated to dryness in iron pots and calcined to burn away the organic matter. The residue is the so-called American *potash*, that is, pot-ashes. Instead of evaporating the aqueous extract to dryness, a purer product can be obtained by evaporating the liquid until the less soluble impurities crystallize out, and finally evaporating the mother liquid to dryness as before. White refined potash is sometimes called *pearl ash*. Potash is also obtained from the residue left after beet sugar has been fermented, and the alcohol removed by distillation. The liquid in the retort is evaporated to dryness, calcined, and extracted with water as before.

The potash found in plants is derived from the soil, and the potash in the soil is one product of the decomposition of rocks which form the earth's crust. The potash which herbivorous animals—*e.g.* sheep—draw from the land is largely exuded as an oily sweat from the skin, and called, after the French, *suint*. The suint accumulates in the wool so that it may

form as much as one-third the weight of raw merino wool. The liquid in which wool is first washed contains most of the suint. This liquid can be evaporated to dryness and heated in iron pots or retorts. Potassium carbonate is extracted from the residue by lixiviation with water as indicated above. Much of the potassium carbonate in commerce is manufactured from potassium sulphate by Leblanc's process, to be described later.

Potassium carbonate exhibits the typical alkaline reaction, for it turns red litmus blue, but, like calcium carbonate, it is a salt, and is formed by the union of the base potassium oxide, K_2O , with carbon dioxide, CO_2 . When treated with an acid, the carbon dioxide is expelled, and another salt is formed. Thus with hydrochloric acid: $K_2CO_3 + 2HCl = 2KCl + CO_2 + H_2O$.

The alkaline lakes of Nevada and South California give sodium carbonate on evaporation. "Natural soda," also called *trona* or *urao*, has been extracted commercially from a number of localities. Owen's lake in California has been estimated to contain between 20 and 40 million tons of alkali carbonate. The deposit at Searle's lake also contains much alkali. Trona has a composition corresponding with $Na_2CO_3 \cdot NaHCO_3 \cdot 2H_2O$. Trona also occurs in Egypt. The deposits at Wadi Natrun in the Libyan district (N. Africa), and at Zoutpan (S. Africa), are worked.

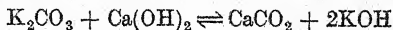
Historical.—While the ashes of land plants furnish potassium carbonate, the ashes of sea plants furnish a similar, but not identical, alkali—sodium carbonate mixed with a little potassium carbonate. Both sodium and potassium carbonates were once included under the Arabian term *alkali*. In order to distinguish these two salts from ammonium carbonate, they were termed *fixed alkalies*, and ammonium carbonate was called *volatile alkali*. H. L. Duhamel du Monceau (1736) first clearly recognized the difference between potash and soda, and the two were distinguished by using the term *vegetable alkali* for potassium carbonate, and *mineral alkali* for sodium carbonate. When M. H. Klaproth (1796) showed that the vegetable alkali occurred in many minerals; the term *potash* (English), or *potasse* (French), was applied to this particular compound. The Germans use the term *kali*—derived from the Arabian term "kali," for ash. Klaproth also proposed to confine the term *natron* to mineral alkali, sodium carbonate. The English equivalent for natron is *soda*, and the French *soude*. In the Middle Ages, the Italian word "soda" was applied to all alkalies, and particularly to an ash used in making glass; and the French applied the term "soude" to the plant glasswort, the ashes of which were used in making glass and soap. The first letter of the word "kali" is used by all chemists as the symbol for potassium, and the first two letters of the word "natron" for sodium.

§ 2. Potassium and Sodium Hydroxides.

Preparation.—Sodium and potassium hydroxides cannot be conveniently made by calcining the corresponding carbonate and digesting the residue with water as in the case of calcium hydroxide, because the two alkali carbonates do not decompose so readily as calcium carbonate. The preparation of the oxides of potassium and sodium is rather difficult and expensive. Hence, although the hydroxide can be made by the action

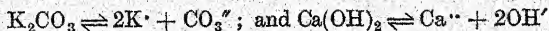
of water on these oxides, it is far more economical to employ other methods of preparation. Two processes are used in the manufacture of these hydroxides. Let potassium hydroxide be taken as a type for both: When calcium hydroxide is added to a boiling solution of potassium carbonate in an iron or silver or nickel vessel, calcium carbonate is precipitated, and potassium hydroxide, KOH, remains in solution: $\text{Ca(OH)}_2 + \text{K}_2\text{CO}_3 = 2\text{KOH} + \text{CaCO}_3$. The clear solution is decanted from the precipitated calcium carbonate, and concentrated by heating it in iron pots. The electrolysis of an aqueous solution of potassium chloride furnishes chlorine gas (*g.v.*), and a solution of potassium hydroxide. The latter is concentrated by evaporation. In Acker's process (*g.v.*) fused chloride is used in place of an aqueous solution.

Theory of preparation—molecular.—The reaction between potassium carbonate and calcium hydroxide has many points of interest. It is best studied in the light of the theory of equilibrium, so useful in the study of chemical reactions generally. The four salts in solution are in equilibrium, and accordingly, the reaction is represented:



An excess of solid calcium hydroxide is supposed to be present at the start so that as fast as calcium hydroxide is removed from the solution by reacting with the potassium carbonate more passes into solution. Thus the concentration of the calcium hydroxide in the solution is kept constant. The solubility of calcium carbonate is very small, and, in consequence, any calcium carbonate in excess of the solubility constant will be precipitated as fast as it is formed. The reaction proceeds steadily from left to right because, all the time, calcium hydroxide steadily passes into solution, and calcium carbonate is steadily precipitated. But the solubility of calcium carbonate steadily increases with increasing concentrations of potassium hydroxide. There is a steady transformation of the potassium carbonate into potassium hydroxide in progress. The concentration of the potassium carbonate is steadily decreasing, while the concentration of the potassium hydroxide is steadily increasing. Consequently, when the potassium hydroxide has attained a certain concentration so much calcium carbonate will be present in the solution that the reaction will cease. Hence the concentration of the potassium carbonate should be such that it is all exhausted before the state of equilibrium is reached. If the concentration of the potassium hydroxide should exceed this critical value the reaction will be reversed, and calcium carbonate will be transformed into calcium hydroxide.

Theory of preparation—ionic.—The explanation offered by the ionic theory runs somewhat as follows: At the start, the solution contains the ions



The solubility product $[\text{Ca}^{++}][\text{CO}_3^{--}]$ is very small, and very much less than the solubility product $[\text{Ca}^{++}][\text{OH}']^2$; consequently, since relatively large proportions of both the ions Ca^{++} from the Ca(OH)_2 , and CO_3^{--} from the K_2CO_3 are present in the solution, calcium carbonate will be precipitated, and will continue being precipitated so long as the potassium carbonate and calcium hydroxide can supply ions CO_3^{--} and Ca^{++} in excess of the solubility product of calcium carbonate. But the calcium hydroxide

furnishes the Ca^{++} ions, and the solubility of $\text{Ca}(\text{OH})_2$ is determined, as we have seen, by the solubility product $[\text{Ca}^{++}][\text{OH}']^2$. With the steady removal of Ca^{++} and CO_3^{--} ions, the concentration of the K^+ and the OH' ions must be continually increasing. By-and-by the concentration of the OH' ions becomes relatively large; this, in virtue of the common ion OH' , reduces the concentration of the Ca^{++} ions required to maintain the solubility product of calcium hydroxide up to its own constant value. Finally, when the concentration of the OH' ions is so great that the concentration of the Ca^{++} ions from the calcium hydroxide is no greater than the concentration of the Ca^{++} ions required to maintain the solubility product of the calcium carbonate at its own characteristic value, the reaction will stop. Hence the concentration of the potassium carbonate in the solution should be so adjusted that this salt is exhausted before the reaction stops.

Properties.—Both hydroxides are white crystalline compounds which rapidly absorb moisture and carbon dioxide from the atmosphere. In aqueous solution, both hydroxides are corrosive bases. Both compounds dissolve in water with the evolution of much heat. Both hydroxides melt easily, and are frequently cast into sticks for convenience in use. Potassium hydroxide becomes anhydrous at 400° , it does not appreciably dissociate although it vaporizes slightly at 600° . Gold is attacked, but not chemically, by the fused hydroxide; nickel forms nickel hydroxide. The latter action is so slow that nickel crucibles and dishes are employed for the evaporation or fusion of caustic alkalis. Gold, silver, and iron vessels are also used. When a small proportion of alkali hydroxide is added to solutions of various metal sulphates, a basic sulphate is usually precipitated, except in the case of magnesium and manganese sulphates. With an excess of the alkali-lye, the product is usually a hydroxide, but not always—e.g. copper and nickel may form a series of basic sulphates. Large quantities of sodium hydroxide—also called *caustic soda*—are used in bleaching, dyeing, and in refining of oils, in the manufacture of hard soap, and of paper. Potassium hydroxide—also called *caustic potash*—is also used for making soft soap.

§ 3. Black's Investigation on the Carbonates of the Alkalies and Alkaline Earths.

Black's experiments are so solid and convincing that it seems to me impossible to form any objection against them.—C. W. SCHEELÉ.

J. Black's "experiments upon magnesia alba, quicklime, and other alkaline substances," published in 1755, first made clear the relations between caustic alkali and mild alkali;¹ that is, between the alkali oxides and alkali carbonates. These relations were not understood by the early chemists. They believed the "mild alkalies" and "mild earths" to be elementary substances; that the causticity of lime was due to the union of "fire-matter" (phlogiston) with the element chalk; and the conversion of mild alkali into caustic alkali with the simultaneous regeneration of chalk by boiling the former with caustic lime, was due simply to

¹ The term "mild alkalies" was formerly applied to what are now called "alkali carbonates."

the transfer of the "fire-matter" from the lime to the mild alkali. Otherwise expressed :

$$\text{Quicklime} = \text{Chalk} + \text{Fire-matter.}$$

Black proved this hypothesis to be untenable.

Black demonstrated experimentally that chalk after ignition neutralized the same quantity of acid as before ignition, but the calcined chalk dissolves in the acid without effervescence, whereas the original chalk lost a gas which he called "fixed air," but which is now called "carbon dioxide." The salts formed by the action of acids on calcined and uncalcined lime are identical in every respect, and the same amount of gas is expelled from chalk whether the chalk be calcined or digested in acids. Further, by weighing the chalk before and after calcination, Black found a *loss*, not a gain, in weight. Thus :

Ordinary chalk	120 grains
Quicklime	68 "
Loss in weight	52 "

Hence, added Black, "we may safely conclude that the volatile matter lost during the calcination is mostly air, and hence calcined lime does not emit air or make an effervescence when mixed with acids." Again, lime becomes caustic owing to the loss of fixed air. Consequently Black proved :

$$\text{Chalk} = \text{Quicklime} + \text{Fixed air.}$$

Hence, quicklime is simpler than chalk or limestone.

On boiling the 68 grains of quicklime obtained in the experiment cited above, with pot-ashes, Black finally obtained 118 grains of a white powder "similar in every trial" to ordinary chalk. The 118 grains of chalk correspond with the 120 grains originally taken within the limits of experimental error. The resulting caustic pot-ashes no longer effervesced with acids, whereas the regenerated chalk did. Hence Black concluded that the pot-ashes were made caustic by the transfer of the gas contained in pot-ashes to the caustic lime. Somewhat similar results were obtained with magnesia and its carbonate as were obtained with lime and its carbonate.

Black thus demonstrated the modern view of the changes which attend the transformation of a mild into a caustic alkali, and proved that these changes are similar to those which occur during the conversion of a mild earth into a quicklime. Black's experiments also made clear the relations between the "mild alkalies" (alkali carbonates); "caustic alkalies" (alkali hydroxides); "mild earths" (carbonates of the alkaline earths); and the "quicklimes" (oxides of the alkaline earths). These interesting experiments by Black involved the use of the balance. Prior to this, chemistry had laboured only with the *qualities* of bodies; Black recognized the importance of *quantities*, and this opened up a new era in science. Soon afterwards, T. Bergman emphasized the importance of the balance as "of singular service in investigating the properties of bodies and in directing those properties to their proper use. All effects are exactly proportioned to their causes; therefore, unless their mutual relations be examined by accurate trials, theory must be lame and imperfect." Black's methods, later on, were extended by Lavoisier in his brilliant work on oxidation and combustion.

§ 4. Metallic Sodium and Potassium.

The key to a knowledge of Nature's methods is in her analogies.—R. OTTO-LENGUI.

When any apparently exceptional or new substance is encountered, the chemist is guided in his treatment of it by analogies which it seems to present with previously known substances.—W. S. JEVONS

Discoveries in science are very often made by following up hints received from analogies. The isolation of the metals potassium and sodium is a good illustration. At the beginning of the 19th century, the so-called alkalis and alkaline earths—magnesia, lime, and potash—were considered to be elementary substances. Lavoisier proved that some things resembling the earths—*e.g.* tin oxide, iron rust, mercuric oxide—could be resolved into two substances, oxygen and a metal. By analogy, Lavoisier inferred that the indifference of the earths for oxygen is due to their being already oxides, and he ventured to predict that the earths would soon cease to be counted among simple substances, and that it would be possible to resolve them into the corresponding metals and oxygen. It was reserved for H. Davy to realize this prophetic conjecture. After it had been shown that the electric current could resolve water and certain other salts into their elements, H. Davy tried if the electric current would work in an analogous manner on caustic soda and caustic potash. As a result, Davy isolated the metal potassium on October 6, 1807, and sodium a few days afterwards. This discovery was soon followed by the isolation of barium, strontium, and calcium. By analogy, it was further inferred that all amorphous powders—alumina, magnesia, etc.—possessing similar properties, were metallic oxides. As a result, when a new earth is now discovered, chemists believe, by faith, that it is the oxide of a metal even in cases where the supposed metal has never been isolated.

Davy exposed a piece of solid potassium hydroxide to the atmosphere for a few seconds so that a conducting film of moisture formed on the surface. The piece of potash was then placed on an insulated disc of platinum connected with the negative pole of a battery, and a platinum wire connected with the positive pole was brought in contact with the upper surface of the potash. Davy adds:

Under these circumstances a vivid action was observed to take place. The potash began to fuse at both its points of electrification. There was a violent effervescence at the upper surface; at the lower, or negative surface, there was no liberation of elastic fluid; but small globules having a high metallic lustre, and being precisely similar in visible character to quicksilver, appeared, some of which burnt with explosion and bright flame, as soon as they were formed, and others remained, and were merely tarnished, and finally covered by a white film which formed on their surface. These globules, numerous experiments soon showed to be the substance I was in search of, and a peculiar inflammable principle, the basis of potash.

Soon after Davy's discovery, J. L. Gay-Lussac and L. J. Thénard (1808) prepared the metal by heating metallic iron with potash at a white heat: $4\text{KOH} + 3\text{Fe} = \text{Fe}_3\text{O}_4 + 2\text{H}_2 + 4\text{K}$. The potassium metal vaporized, and condensed in a copper receiver containing rock oil to prevent the oxidation of the metal. Later on, M. Curaudau (1808) substituted charcoal in place of iron as reducing agent; and later still, H. Y. Castner

(1888) used iron carbide— FeC_2 . Calcium carbide, CaC_2 , has also been used for the same purpose: $6\text{KOH} + 2\text{CaC}_2 = 6\text{K} + 2\text{CaO} + 4\text{CO} + 3\text{H}_2$. The element is also obtained by the action of carbon on potassium carbonate at a high temperature: $\text{K}_2\text{CO}_3 + 2\text{C} = 2\text{K} + 3\text{CO}$, but if the vapour of the potassium be not rapidly cooled, a black explosive compound with the carbon monoxide, CO , is said to be formed—potassium carbonyl, $\text{K}_2(\text{CO})_2$, or K.CO.CO.K . In modern works, where cheap electrical energy is available, modifications of Davy's original process—electrolysis of fused sodium hydroxide—are used for preparing sodium industrially, e.g. H. Y. Castner's electrolytic process (1890). Sodium chloride mixed with powdered lead, and heated red hot in a closed retort is said to give metallic sodium: $2\text{NaCl} + \text{Pb} \rightarrow 2\text{Na} + \text{PbCl}_2$. By heating alkali chlorides with metallic calcium, in vacuo, caesium and rubidium have been prepared, but lithium could not be so obtained because it forms an alloy with calcium.

H. Y. Castner's electrolytic process for sodium.—The sodium hydroxide, contained in an iron pot set in brickwork, is melted by means

of a ring of gas jets placed underneath; and kept about 20° above the melting point (310°) of sodium hydroxide. The cathode, *H*, rises through the bottom of the iron pot, *A*, Fig. 147, and is maintained in position by a cake, *K*, of solid sodium hydroxide in the lower part of the pot. The anodes, *F*, several in number, are suspended around the cathodes from above. A cylindrical vessel, *ND*, floats in the

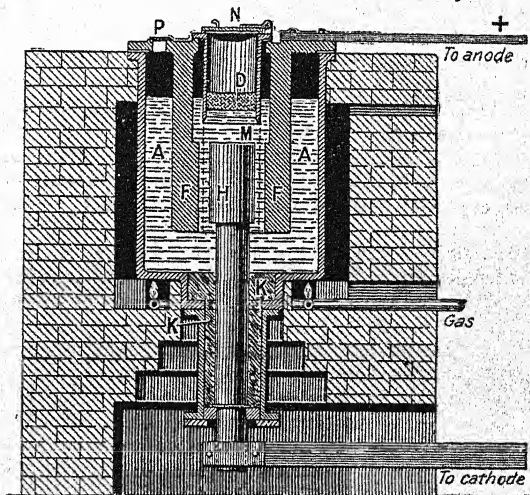


FIG. 147.—Castner's Electrolytic Process for Sodium.

fused alkali above the cathode, and the sodium and hydrogen liberated at the cathode collect under this cylinder. The hydrogen escapes through the cover, and the atmosphere of hydrogen in the cylinder protects the sodium from oxidation. A nickel wire cage, *M*, separates the anode, *F*, from the cathode, *H*. From time to time the sodium, *D*, is skimmed off by means of a perforated ladle which retains the liquid metal, but allows the molten hydroxide to flow back. The oxygen liberated at the anode escapes *via* the vent *P*. Hydrogen is a by-product in Castner's process, and if there be no commercial demand for this gas, its production represents so much wasted energy. Attention has been previously directed to the preparation of sodium by the electrolysis of fused sodium chloride whereby sodium is produced at one electrode, and chlorine at the other;

but the production of sodium by the electrolysis of the fused chloride has not been very successful commercially.

Properties.—Both sodium and potassium are silvery white lustrous metals which tarnish at once when exposed to the air owing to the formation of a film of oxide. These metals are therefore usually preserved in well-stopped vessels, or in a liquid containing no oxygen, *e.g.* naphtha, or petroleum. In thin layers on glass, potassium appears bluish-violet; sodium, yellowish-brown. The vapour of potassium at a red heat appears to be violet, and sodium colourless in thin layers, purple in thick layers. The metals are lighter than water, and at ordinary temperatures they are soft enough to be moulded between the fingers. Sodium melts at 97.6° , potassium at 62.4° ; sodium boils at 877° , potassium at 758° . Potassium is rather more chemically active than sodium. Both metals react with water to form hydroxides: $2K + 2H_2O = 2KOH + H_2$. The heat of the reaction with potassium suffices to ignite the hydrogen; with sodium, the hydrogen ignites if the water is warmed. The flame of hydrogen is coloured by the vapours of the respective metals—potassium, lavender-blue; sodium, daffodil-yellow. Cæsium begins to give off hydrogen in contact with ice at -116° ; rubidium at -108° ; potassium at -105° ; and sodium at -98° . Sodium and potassium dissolve in liquid ammonia, giving blue solutions; with potassium the solution is indigo-blue. Heated in an atmosphere of carbon dioxide, free carbon and a carbonate of the metal are formed. When heated in air, sodium and potassium burn, each metal forming a mixture of oxides. Perfectly dry air or oxygen has no appreciable effect upon the dry metals. Potassium with carbon monoxide forms the explosive compound previously mentioned; with the halogens, the metals take fire, forming the corresponding halides. When heated at about 350° in hydrogen, white acicular crystalline hydrides are formed—KH and NaH. The crystals of sodium hydride resemble those of the alkali halides, and have not the characteristics of a metallic alloy. Distilled sodium gives off about twenty times its volume of occluded hydrogen when heated in vacuo, and the same gas is evolved when sodium is dissolved in mercury.

Uses.—An alloy of potassium and sodium, liquid at ordinary temperatures, is used for some high-temperature thermometers above the boiling point of mercury. Sodium is used in the manufacture of sodium cyanide and sodium peroxide, in drying oils, and in the manufacture of organic compounds.

Atomic and molecular weights.—The combining weight of sodium determined by the analysis of sodium salts—sodium chloride, etc.—lies between 23.00 and 23.17; the best representative value is probably 23. The atomic weight by Dulong and Petit's method of approximation ($6.4 \div \text{specific heat} = \text{atomic weight}$) is $6.4 \div 0.283 = 22.6$. Consequently, 23 is taken to be the atomic weight of sodium. The vapour density of sodium is 25.7, which corresponds with a one-atom molecule. Similarly, analyses of potassium chloride and other potassium salts show that the combining weight of potassium lies between 38.67 and 39.33; the best representative value is supposed to be 39.10; and, the specific heat being 0.166, Dulong and Petit's method of approximation furnishes $6.4 \div 0.166 = 38.56$, corresponding with the atomic weight 39.10. The vapour density 42.34 at 1040° corresponds with a one-atom molecule. Freezing-point determina-

tions of solutions of the metals in mercury show that these elements have a one-atom molecule when dissolved in mercury. The vapour densities of potassium and sodium chlorides at 2000° show that here again the atomic weights of potassium and sodium are probably 39.10 and 23 respectively.

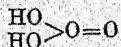
Occurrence.—Potassium and sodium only occur in nature combined with other elements, but the compounds are widely distributed, being present in many silicate rocks, etc. Potassium occurs as sylvine, carnalite, kainite, etc., in the Stassfurt deposits (*q.v.*). Sodium is present in sea-water, etc., and it occurs as rock salt (*q.v.*) and as Chili saltpetre (*q.v.*).

§ 5. Potassium and Sodium Monoxides and Peroxides.

Sodium and potassium monoxides— Na_2O and K_2O .—The monoxides of the alkali metals are made by heating the metals in a limited supply of dry air at a temperature below 180°, and removing the excess of metal by distillation in *vacuo*. The monoxides so obtained are not pure, being probably a mixture of monoxide, say Na_2O , and peroxide, Na_2O_2 . The monoxides have also been made by heating the metal with the corresponding nitrates: $2\text{NaNO}_3 + 10\text{Na} = 6\text{Na}_2\text{O} + \text{N}_2$. The two oxides combine with water to form the hydroxides with the evolution of a considerable amount of heat. Lithium oxide, however, does not evolve so much heat.

Sodium peroxide, Na_2O_2 .—This oxide is formed when sodium is burnt in a stream of oxygen, or when the metal is heated in aluminium trays in iron tubes at about 300°. The product contains about 95 per cent. of sodium peroxide. Pure sodium peroxide is white, but the commercial product is usually tinged yellow. The peroxide is stable in dry air at ordinary temperatures; and in moist air, or in the presence of water, it is decomposed with the evolution of oxygen and the formation of sodium hydroxide: $2\text{Na}_2\text{O}_2 + 2\text{H}_2\text{O} = 4\text{NaOH} + \text{O}_2$; but if the temperature be kept low, hydrogen peroxide is formed: $\text{Na}_2\text{O}_2 + 2\text{H}_2\text{O} = 2\text{NaOH} + \text{H}_2\text{O}_2$. A hydrate, $\text{Na}_2\text{O}_2 \cdot 8\text{H}_2\text{O}$, has been prepared. Sodium peroxide is a powerful oxidizing agent, and it is used in the laboratory for oxidizing purposes and also for decomposing silicate rocks prior to analysis. Sodium peroxide is also used in straw bleaching, etc. The heats evolved when potassium and sodium monoxides are converted into the corresponding peroxides are respectively 22 Cals. and 19 Cals.

Potassium tetroxide, K_2O_4 .—This oxide is formed as a chrome-yellow powder when metallic potassium is melted in an atmosphere of nitrogen, and the nitrogen is gradually displaced by air or oxygen. All traces of moisture must be avoided. It is also formed when potassium is heated in nitrous oxide, N_2O . If the temperature be suitably regulated, K_2O_2 is said to be formed, but there is some doubt about this. When treated with water, potassium tetroxide forms potassium hydroxide, hydrogen peroxide, and oxygen; when heated with carbon monoxide, potassium carbonate and oxygen are formed: $\text{K}_2\text{O}_4 + \text{CO} = \text{K}_2\text{CO}_3 + \text{O}_2$. A. von Bayer and V. Villiger consider that potassium and rubidium tetroxides are salts of what they called ozonic acid, H_2O_4 , or



On heating the tetroxide to 480° in *vacuo*, R. de Forcrand obtained what

he considered to be potassium trioxide, K_2O_3 , which gave potassium sulphate, hydrogen dioxide, and oxygen when dissolved in dilute sulphuric acid. The heat of formation of sodium trioxide is about 126 Cals., a number nearly the same as that for the corresponding potassium compound.

§ 6. Spectrum Analysis.

If there ever was a flank movement of Nature by which she has been compelled to surrender a part of her secrets, it was the discovery of the spectro-scope, which enables us to peer into the very heart of Nature.—R. C. KEMPZIE.

Near the beginning of the 17th century Kepler's curiosity was roused by the play of colours which he observed when a beam of sunlight was

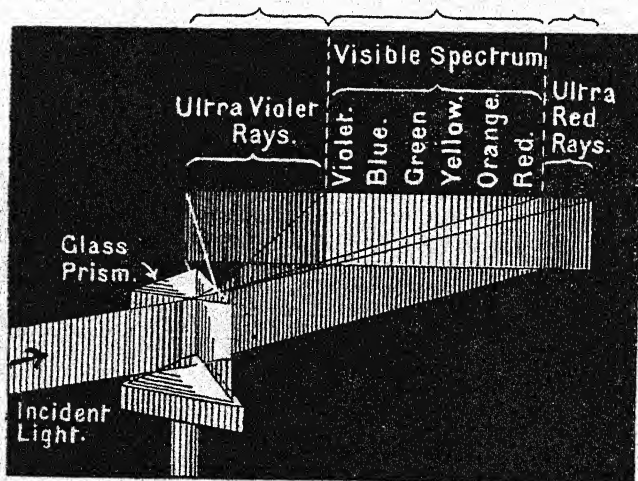


FIG. 148.—Newton's Experiment.

passed through a glass prism, and nearly half a century later, Isaac Newton (1666) proved experimentally that a beam of sunlight is composed of light rays of various colours perfectly blended and ranging from red through orange, yellow, green, and blue to violet. This Newton did by passing the beam of sunlight through a glass prism, and projecting the beam on to a screen, Fig. 148. The violet, green, and blue rays are bent more in passing through the prism than the yellow, orange, and red rays. The beam of light after passing through the prism thus appears on the screen as an unbroken band of colours, which is called a continuous spectrum. Any beam of "white" light can be used in place of sunlight; for instance, the light from an incandescent solid such as the limelight, Welsbach's mantle, incandescent carbon, etc. Newton admitted the light through a circular aperture. W. H. Wollaston (1802) admitted the light through a narrow slit and noticed that the beam of sunlight really furnished a spectrum which is crossed by a large number of dark lines—some sharp

and well defined, others more or less faint and nebulous. J. von Fraunhofer (1814-15) happened to look at Wollaston's spectrum through a small telescope and discovered hundreds of these lines. He carefully mapped the relative positions of a number of these dark lines, and accordingly they are now generally called **Fraunhofer's lines**. It was soon found that incandescent vapours and gases furnish a discontinuous or line spectrum, that is, a spectrum composed of a few bright lines instead of a continuous band. The line spectra of some elements are comparatively simple, for they display but a few clear distinct coloured lines—*e.g.* sodium a yellow line; thallium a green line; indium a blue and an indigo-blue line. Other elements have complex spectra containing numerous lines of varying intensity—*e.g.* barium, strontium, iron. The spectra of some elements, though complex, are easily recognized by the prominence and position of certain lines—*e.g.* the dark red line of rubidium, the blue line of caesium, etc. Brewster, and later Herschell, found that when a coloured glass is placed between the source of light and the slit, the spectrum disappeared, all except the light having the same colour as the glass. Brewster felt that the cause of the quenching of the lines in the solar spectrum must be absorption; but he could get no further. The human mind, as E. Larkin expressed it, was now groping in darkness in the presence of one of the chief mysteries of nature. The cause of the dark lines in the solar spectrum was quite inexplicable. Fraunhofer died without knowing their cause. Let us therefore follow the path of the physicists, who, with consummate skill, discovered the mystery behind the dark lines in the solar spectrum.

It was provisionally assumed that "Fraunhofer's lines are due to the lack of certain shades of colour in the spectrum of sunlight." This plausible hypothesis was tested by L. Foucault in 1849. He arranged an experiment in which a ray of sunlight was directed by means of lenses on to the glowing gases between the poles of an arc light which alone gave a spectrum with two yellow lines very prominent. The mixed light was passed through a prism. Foucault expected the lacuna in the solar spectrum corresponding with the yellow lines of the glowing gases in the arc light would be filled by the latter; and that the yellow portion of the solar spectrum would be continuous. To his astonishment, the reversed or dark lines of the solar spectrum corresponding with the bright yellow lines were more pronounced than before. Hence the preliminary hypothesis cannot be true.

Foucault then suggested the hypothesis that "the incandescent gases in the arc light have the power of absorbing the yellow from sunlight, as well as of emitting yellow light. The increase in the darkness of these lines is due to this absorbing power being greater than the emissive power." Foucault focussed the light from an incandescent carbon—which by itself gives a continuous spectrum—on to the incandescent gases between the carbon poles. Instead of getting the continuous spectrum of the incandescent carbon with yellow lines enhanced by the spectrum of the gases between the carbon poles, Foucault obtained a spectrum with the dark lines in the yellow portion. This experiment supported his hypothesis. *The same kind of light which a gas gives out when heated will be arrested if the attempt be made to pass the light through the gas.* Had Foucault known that the yellow lines were due to sodium he would have recognized the origin of Fraunhofer's lines. This interpretation was made by G. Stokes,

1852. In 1859 R. Bunsen and G. Kirchhoff definitely settled the question. "This discovery," said E. Larkin, "is equal in importance to that of the law of gravitation, for the one 'weighed' the universe, and the other tells us of what it is made." The subsequent history of spectrum analysis is but an illustration of the fact that once the right explanation of a phenomenon is found the facts seem to arrange themselves about the theory as naturally as the particles of a salt in a solution aggregate about the enlarging nucleus of a crystal.

Bunsen and Kirchhoff subsequently proved that, under similar conditions of temperature and pressure, every element produces its own characteristic spectrum; and reciprocally, the presence of the vapour of an element can be inferred with certainty when the characteristic lines are present. The spectrum of an element is so definite and characteristic that it has been proposed to use this as part of the definition of an element. For instance, A. de Gramont proposed that no new substance be recognized as an element until its spark spectrum has been measured and shown to be different from that of every other known form of matter. The spectrum of the incandescent vapour of a mixture of elements contains all the lines characteristic of each element in the mixture, and consequently it is possible to recognize each and all of them by measuring the position of the bright lines and comparing the lines with those of known elements. This method of detecting elements is called spectrum analysis. Bunsen and Kirchhoff were able to prove that the dark Fraunhofer's lines are due to the rays of light from an incandescent solid passing through vapours of various elements. For the characteristic bright lines of, say, sodium are changed into dark lines when the white light from an incandescent solid is allowed to pass through the vapour of this element. Hence, since the bright lines of the spectrum of a vapour are changed into dark lines, and the dark parts are changed into bright parts by the continuous spectrum of white light, it seems as if the whole spectrum is reversed with respect to its illumination. It is accordingly called the phenomenon of the reversed spectrum. Hence, it was inferred that Fraunhofer's lines are due to the rays of light from an incandescent sun passing through the sun's atmosphere, and consequently the vapours of the elements whose line spectra correspond with the dark lines of the solar spectrum must be present in the sun's atmosphere, and

The spectrum catches tokens from his light
Of elemental kinship with the earth.—LANE.

By the aid of spectrum analysis, therefore, it has been possible to deduce the presence of a large number of known elements—some 34—in the sun's atmosphere from the coincidence of the bright lines furnished by elements in the laboratory with the dark lines in the solar spectrum. The halogen elements, nitrogen, oxygen, gold, mercury, and a few other elements, have not been detected in the sun. The spectrum of an unknown element—helium—was observed in the solar spectrum some thirty years before the corresponding element was discovered in the earth. The light from the fixed stars furnishes results similar to those obtained with sunlight. The lines of hydrogen, helium, carbon, magnesium, calcium, and iron have been detected in nebulae; and hydrogen and hydrocarbons have been recognized in comets. "The spectral lines," continued E. Larkin,

"can thus be said to make up the alphabet of the universe, and spell out messages from our sun and other suns; these can be read by an expert as well as the Morse dots and dashes of the telegraphic code."

The spectroscope.—Quite a large number of instruments have been devised for the examination of the spectra of different substances. The so-called direct vision spectroscopes have a slit at one end to admit the light under examination. The slit can be narrowed or widened by turning a suitable screw. The beam of light passes through a lens which keeps them parallel; after passing from the slit through a prism, they are directed to the eye-piece of a small telescope where the spectrum is examined.

The spectra of solids and liquids.—Solids and liquids must be vaporized before their spectra can be examined. It is often sufficient to introduce salts of the elements into a non-luminous Bunsen's flame by means of a clean platinum wire. The flame spectrum is then examined by means of the spectroscope. This method is satisfactory provided there is plenty of material, and the salt volatilizes and dissociates in the flame. If the substance is not volatilized in the Bunsen's flame, or if but a minute quantity of the substance is available, it is best to use a spark spectrum obtained in the following manner—due to Delachanel and Mermet: A test-tube, Fig. 149, has a platinum terminal fused into the bottom to serve as one electrode; a cork in the mouth of the test-tube supports a glass tube pierced with a platinum wire to serve as the other electrode. The height of the solution in the test-tube is arranged so that a certain portion rises by capillarity and fills a central cavity in the glass tube holding the lower electrode. When the cup is full, every time a spark passes a very minute portion of the solution is vaporized, no material is lost, and the uniformity of the volatilization with each spark permits continuous observations.

The spectra of gases.—The spectrum of a gas is obtained by sealing the gas in a tube, Fig. 150, while the gas is under a reduced pressure. Each end of the tube has a platinum wire sealed into it, and these wires are put in communication with an induction coil. When a series of sparks are passed through the gas, the gas becomes incandescent, particularly in the narrowed portion of the tube, and the characteristic spectrum for that particular gas is obtained.

It might be added that the spectrum of a gas depends partly upon the *temperature* and partly upon the *pressure* under which the gas is confined. New lines are added, and old lines are intensified—enhanced lines—when the temperature of a gas is raised—e.g. the lithium spectrum in the electric arc shows a blue line which is not visible in the spectrum obtained by heating with the gas-flame. Sodium in a Bunsen's flame, for instance, gives one well-defined yellow line which is really compounded of two yellow

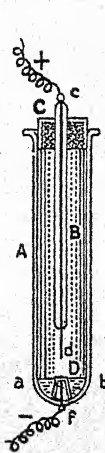


FIG. 149.—
Spectrum Tube
Liquids.

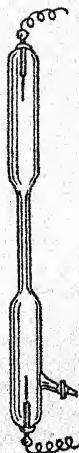


FIG. 150.—
Spectrum Tube
Gases.

lines; but at higher temperatures, three other pairs of lines make their appearance. The difference in the high and low temperature spectra is sometimes so marked that in the early days of spectroscopy, J. Plücker and W. Hittorf (1865) wrote: "There is a certain number of elementary substances which when differently treated furnishes two kinds of spectra of quite different character, not having any line or band in common." We know now that the changes in the spectrum of an element with variations of temperature are themselves characteristic, so that it is possible to deduce the temperature from the corresponding spectrum of a hot body. In a general way, it may be said that the spectra of incandescent elements become more complex as the temperature is raised. In the case of compound bodies, too, E. Mitscherlich showed that they may have a characteristic spectrum in which the lines of the constituent elements can be detected provided the temperature is high enough to decompose the incandescent compound. At small pressures, spectra with broad bands are obtained. As the pressure is increased, "new series of lines arise which existed only in germ at lower pressures"; the bands give way to lines, and finally a continuous band of light is obtained corresponding with the spectrum of an incandescent solid. *The character of the spectral lines thus depends upon the pressure and temperature under which the gas or vapour is examined.* Under great pressures, all incandescent solids, liquids, and gases emit a continuous spectrum.

Absorption spectra.—Some substances absorb certain colours and transmit others. If coloured solutions are examined by the spectroscope illuminated by the light of an incandescent solid transmitted through the solution, a series of dark bands or lines called absorption spectra are obtained, in contrast with *emission spectra* just described where the light emitted by the hot body is examined. A number of solutions have very characteristic absorption spectra—*e.g.* blood, didymium chloride, copper sulphate, potassium chromate, potassium permanganate, potassium dichromate, etc.

The delicacy of spectrum analysis.—The spectroscope is one of the most delicate means of detecting many substances, and it enables elements to be recognized with certainty when present in quantities far too small to produce an appreciable effect upon the most sensitive reagents known. Thus 1 c.c. of air contains approximately 0.0001 c.c. of neon, and the neon in $\frac{1}{20}$ c.c. of ordinary air, that is, 0.000005 c.c. of neon has been detected by means of the spectroscope. By means of the spectroscope also it is possible to detect the presence of 0.00006 milligram of strontium and of calcium; 0.00001 milligram of lithium; and 0.0000003 milligram of sodium. It is not likely that rubidium and caesium would have been discovered so soon had it not been for their striking spectra. Thallium, indium, and gallium were also discovered by the aid of the spectroscope.

§ 7. Lithium, Rubidium, and Caesium.

Lithium, Li.—The three elements—lithium, rubidium, and caesium—are related to potassium and sodium. Lithium was discovered by A. Arfvedson, in 1817, while analyzing the mineral petalite. He found the sodium to be contaminated with an alkali which, unlike potassium, was not removed by treatment with tartaric acid, and, unlike sodium, formed a sparingly soluble carbonate. The name lithium is derived from the

Greek *λίθος* (litheos), stony, because it was believed, at the time of its discovery, that its presence was confined to the mineral kingdom. Lithium is widely distributed in small quantities. It occurs in a number of minerals: *lepidolite*, or lithia mica, contains up to about 6 per cent. of lithia; *spodumene*— $\text{Li}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 4\text{SiO}_2$ —up to about 6 per cent.; *amblygonite*, $\text{Li}(\text{AlF})\text{PO}_4$, has 7–10 per cent. of lithia, and *petalite* has up to about 3 per cent. of lithia—*castorite* is a variety of petalite. Lithium has been detected in sea-water and in most spring and river waters. W. A. Miller (1864) has reported 0.37 gram per litre in the water of a mine at Redruth (Cornwall). Lithium has also been detected in the ash of many plants—tobacco, sugar cane, etc.; in the ash of milk, blood, etc.; and also in a number of meteorites.

To extract lithium, the powdered mineral is calcined with a mixture of ammonium chloride and calcium carbonate. The aqueous extract is treated with hydrochloric acid and evaporated to dryness. The lithium chloride is extracted with amyl alcohol or pyridine in which the lithium chloride is fairly soluble. The solvent is then removed by evaporation; lithium chloride remains. R. Bunsen and A. Matthiessen prepared relatively large quantities of metallic lithium, in 1855, by the electrolysis of the fused chloride; many others appear to have obtained the metal before 1855, but in quantities too small for examination.

When the solution of the chloride is treated with ammonium carbonate, lithium carbonate, Li_2CO_3 , is precipitated. Lithium carbonate, unlike the other alkaline carbonates, is decomposed at a high temperature, and in this respect resembles the carbonates of the alkaline earths. The metal is made by the electrolysis of the fused chloride. The metal resembles sodium and potassium. It decomposes water, but the hydrogen does not ignite even if the water be boiling. When heated in oxygen, the metal forms lithium monoxide, or lithia, Li_2O , but it does not readily form a higher oxide. The monoxide can be prepared by the dissociation of the carbonate at 780° – 800° in a current of dry hydrogen; and also by the dehydration of the hydroxide under similar conditions. Lithium metal unites rapidly with hydrogen at a red heat, forming lithium hydride, LiH ; with nitrogen it forms lithium nitride, Li_3N . Lithium carbonate and lithium phosphate, unlike the other alkaline carbonates and phosphates, are but sparingly soluble in water. Lithium salts are sometimes used in the treatment of diseases due to uric acid poisoning.

Cæsium and rubidium.—R. Bunsen and G. Kirchhoff (1860), while investigating the mineral waters of Dürkheim (Palatinate), evaporated down 40 tons of the water, and removed the alkaline earths, and lithia with ammonium carbonate. The filtrate showed the spectral lines of sodium, potassium, and lithium, “and besides these, two splendid blue lines” near to the blue strontium line. Bunsen and Kirchhoff add:

As no elementary body produces two blue lines in this portion of the spectrum, we may consider the existence of this hitherto unknown alkaline element was thus placed beyond doubt. The facility with which a few thousandths of a milligram of this body may be recognized by the bright blue light of its incandescent vapour, even when mixed with large quantities of more common alkalis, has induced us to propose for it the name *cæsius* (and the symbol Cs), derived from the Latin *cæsius*, used to designate the blue of the clear sky.

Again, on extracting the alkalis from lepidolite (Saxony) and washing the precipitate obtained by treating the solution of the alkalis with

hydrochloroplatinic acid with boiling water a number of times, the residue finally gives "two splendid violet lines" between those due to strontium and to potassium, as well as a number of other lines in the red, yellow, and green portions of the spectrum. Bunsen and Kirchhoff say:

None of these lines belong to any previously known body. Amongst them are two which are particularly remarkable in lying beyond Fraunhofer's line in the outermost portion of the red solar spectrum. Hence we propose for this new metal the name rubidium (and the symbol Rb), from the Latin *rubidus*, which was used to express the darkest red colour.

Compounds of the two elements are so like those of potassium that they cannot be distinguished from that element by the ordinary tests. The only satisfactory means of detecting the two elements is by spectrum analysis. Rubidium and caesium occur together in lepidolite, in carnallite, and in some porphyries. Lepidolite does not contain a quarter per cent. of rubidia, while carnallite may contain up to 4 per cent. of rubidia. They also occur in many mineral waters, in sea-water, in the ashes of plants. The mineral *pollux*—caesium aluminosilicate—contains the equivalent of some 34 per cent. of caesia. Otherwise both elements occur in very small quantities, and somewhat widely diffused in nature. If present in a mineral, both elements will be found with the alkalis after separating the other elements with ammonium carbonate, etc. The residue when treated with hydrochloroplatinic acid furnishes the chloroplatinates of the alkalis. The lithium and sodium salts can be removed by washing with alcohol. Potassium, rubidium, and caesium salts, unlike sodium and lithium salts, give white crystalline precipitates with perchloric or tartaric acid; gelatinous precipitates with hydrofluosilicic acid; and yellow precipitates with hydrochloroplatinic acid, and with sodium nitratocobaltate. The three remaining elements—potassium, rubidium, and caesium—can be separated by taking advantage of the difference in the solubilities of their alums (see "Alums"): 100 c.c. of water at 17° dissolve 0.62 gram of caesium alum; 2.27 grams of rubidium alum; and 13.5 grams of potassium alum. In passing, it may be mentioned that the same amount of water dissolves 51 grams of sodium alum. The chlorostannates— Rb_2SnCl_4 and Cs_2SnCl_6 —are much less soluble in water than the corresponding potassium salt, and hence rubidium and caesium can be separated from potassium by converting the salts into alums or chlorostannates. Caesium can be separated from rubidium by treatment with antimony trichloride. The caesium salt $2\text{CsCl}\cdot\text{SbCl}_3$ is precipitated, while the corresponding rubidium salt is soluble in water.

In connection with caesium, it is interesting to note that C. F. Plattner, in 1846, was not able to make his analysis of the mineral *pollux* (from Elba) add up to 100 per cent. He sought in vain for the missing element. In spite of adverse criticism, Plattner insisted upon the general accuracy of his work; and it demanded some confidence to publish the result. An unscrupulous man might have cooked his result, and a man imbued with less faith in his work would have suppressed the result as absurd. Not so Plattner; he waited. After Bunsen and Kirchhoff had discovered caesium, F. Pisani (1864) showed that Plattner had mistaken caesium (atomic weight 132.8) for potassium (atomic weight 39.1). By making the corresponding correction, Plattner's analysis was found vindicated, for it was found to be quite satisfactory in the light of the facts known when the analysis was made. To make this quite clear, suppose that 5 grams of a compound, supposed to be potassium chloride, are obtained. This will be multiplied by 0.631 to get the equivalent amount, 3.16 grams of K_2O ; but if the compound be CsCl , not KCl , then the weight must be multiplied by 0.835 to get the corresponding amount,

4.18 grams of Cs_2O . The analysis would thus appear to be 4.18 less 3.16, that is 1.02 grams too low if the 5 grams of caesium chloride were mistaken for potassium chloride. This is a remarkable tribute to the accuracy of Plattner's analysis.

Metallic rubidium is prepared by heating an intimate mixture of the carbonate with finely divided carbon; metallic caesium is prepared by heating the hydroxide with magnesium, or by electrolyzing a fused mixture of caesium and barium cyanides. The barium cyanide is added to make the mixture more fusible. Better yields of both metals have been obtained by heating the chlorides with one-fourth their weight of calcium in exhausted tubes of hard glass. The metals, their oxides and their salts, are closely kin to the salts of potassium and sodium. The tendency of rubidium and caesium to form polyhalides is characteristic.

§ 8. The Relations between the Alkali Metals.

The five elements, lithium, sodium, potassium, rubidium, and caesium, called the alkali metals, exhibit an interesting gradation in the properties of the elements and their compounds in accord with the increase in their atomic weights, from member to member, in passing from lithium to caesium. The metals are silvery white, soft enough to be cut with a knife, rapidly tarnish in air, and decompose water at ordinary temperatures. The lowest temperature at which the action of the different metals on water can be detected is -98° for sodium, -105° for potassium, -108° for rubidium, and -116° for caesium. The elements are all univalent, and manifest a remarkable affinity for oxygen; caesium and rubidium ignite spontaneously if placed in dry oxygen at the room temperature. Sodium and lithium, though compatible with the other members of the family, have feebler affinities. The chemical activity of the alkali metals appears to increase steadily in passing from lithium to caesium. The gradation in the physical properties is illustrated in Table XXIII.

TABLE XXIII.—PHYSICAL PROPERTIES OF THE ALKALI METALS.

	Lithium.	Sodium.	Potassium.	Rubidium.	Cæsium.
Atomic weight . .	6.94	23.0	39.10	85.45	132.81
Specific gravity . .	0.59	0.9723	0.859	1.525	1.903
Atomic volume . .	11.7	23.5	44.4	55.8	71.0
Melting point . .	180.1°	97.6°	62.4°	38.5°	25.0°
Boiling point . .	$+1400^\circ$	877°	758°	696°	670°
Specific heat at 0° .	0.941	0.2811	0.1728	0.0802	0.0522
Coefficient expansion	—	0.000274	0.000282	0.000338	0.000345
Heat of fusion (cals.)	—	27.21	14.67	6.144	3.766
Increased volume on fusion (per cent.) .	1.51	2.03	2.44	—	—
Heat of conversion R_2O to R_2O_2 (Cals.)	7.91	19.03	22.0	24.1	25.0
Lowest temp. of reaction with water .	—	-98°	-105°	-108°	-116°

The elements have a remarkably low specific gravity, and a high atomic volume (*q.v.*). The oxides and hydroxides are markedly basic; they do not exhibit acidic qualities. The physical properties of the salts—solubility in water, molecular volume, optical properties, and the variation in

the form of the crystals show the same order of variation as the atomic weights of the elements. Lithium differs in many respects from the other members of the family. The salts of the alkali metals—nitrates, chlorides, sulphides, sulphates, phosphates, carbonates, etc.—are nearly all soluble in water, although lithium, carbonate, phosphate, and fluoride are very much less soluble than the corresponding salts of the other members. And in this respect, lithium resembles the members of the calcium family, and it thus forms a connecting or bridge element between the alkalis and alkaline earths. The alkali sulphates form isomorphous characteristic alums (*q.v.*), but lithium alum appears to be so soluble that it has not yet been crystallized. The modes of crystallization of sodium and potassium sulphates and carbonates are worth noting. Lithium carbonate is sparingly soluble in water, sodium carbonate is not deliquescent, the others are. The salts of sodium and lithium form stable hydrates with water, whereas potassium, rubidium, and caesium salts are nearly all anhydrous. Sodium resembles lithium in the solubility of its chloroplatinate, acid tartrate, and alum so much so that the alkali metals are sometimes divided into two classes: (1) those with sparingly soluble chloroplatinates—*viz.* potassium, rubidium, and caesium; and (2) those with soluble chloroplatinates—*viz.* sodium and lithium.

§ 9. Atomic Volume.

A connection frequently exists between physical and chemical properties, and in many cases recourse must be had to both for an explanation of a phenomenon to which they may each contribute.—C. L. BERTHOLET (1803).

The quotient obtained by dividing the atomic weight of an element by its specific gravity in the solid condition is called the atomic volume

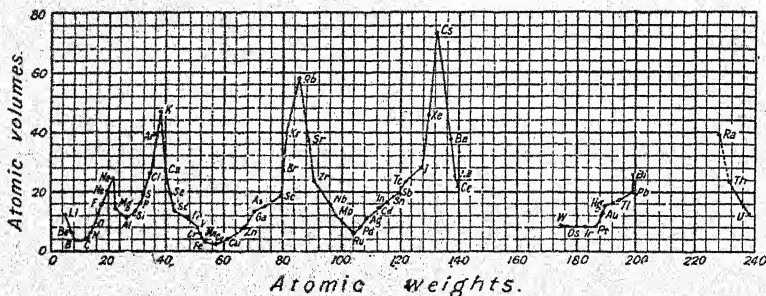


FIG. 151.—Relation between Atomic Volumes and Atomic Weights.

of the element. Consequently, the atomic volume represents the number of cubic centimetres occupied by an amount of the element equal to its atomic weight expressed in grams. The magnitude of the atomic volume thus corresponds with the looseness of texture or porosity, so to speak, of the solid element. Curiously enough, when the atomic volumes of the elements are plotted with the atomic weights, a periodic curve showing a number of maximum and minimum points is obtained, as illustrated in Fig. 151. The atomic volume, for instance, decreases in passing from lithium to boron, after which it increases through carbon,

oxygen, and fluorine to sodium, when it again decreases through magnesium down to aluminium, and thence increases to potassium. Thus the curve passes down and up time and again. A similar curve was obtained in Fig. 114 for the heats of formation of the chlorides. Certain portions of the curve are incomplete owing to the lack of data.

The elements boron, aluminium, cobalt and nickel, rhodium, etc., occupy the troughs of the curve, while the alkali metals occupy the crests of the curve, thus corresponding with the fact that these elements have the largest atomic volumes, or the largest spaces between the atoms. Presumably, the spaces between the atoms of these solid elements are relatively large compared with the size of the atoms themselves. This is often taken to mean that the constituent particles of these elements approximate more nearly to the condition of the particles of a gas than other elements with small atomic volumes. According to D. I. Mendeléeff, the chemical activity of the alkali metals is due to this circumstance; and this assumption is in agreement with the observed increase in the chemical activity of these elements in passing from lithium to caesium.

The atoms of a solid probably do not touch one another, and the volume of a solid thus includes (1) the size of the atoms as well as (2) the spaces between the atoms of the molecule, and (3) the spaces between the molecules. It is not at present possible to distinguish clearly between the effects of these factors.

Questions.

1. What is meant by a normal solution of an acid? How many c.c. of a normal solution of hydrochloric acid would be required to neutralize 1.5 grammes of potassium hydroxide?—*Aberdeen Univ.*

2. Define the following terms: (1) Equivalent of an element, (2) equivalent of a compound, (3) gram-atom, (4) gram-molecule, (5) normal solution.—*Princeton University, U.S.A.*

3. What is meant by the atomic volume of a substance? How is it determined?—*Science and Art Dept.*

4. From what minerals are the salts of potassium prepared and where do they occur? How is potassium nitrate prepared from potassium chloride? How would you show the presence of potassium in a mixture of calcium, sodium, and potassium chlorides?—*Aberdeen Univ.*

5. Describe two or three cases of chemical change in which the influence of "mass" is evident as a factor in the operation.—*London Univ.*

6. Give an account of the chief forms in which calcium carbonate is found. Ten grams of quicklime are slaked, mixed to a paste with water, and added to an excess of sodium carbonate; how much caustic soda is formed ($\text{Ca} = 40$; $\text{Na} = 23$)?—*Owens Coll.*

7. What is a spectrum? Describe the appearance of any spectrum which you have seen.—*London Univ.*

8. Name two compounds containing lithium and describe the preparation of lithium chloride from one of them. In what respects do lithium compounds resemble or differ from corresponding derivatives of other alkali metals?—*Board of Educ.*

9. From your knowledge of the characteristics of opposing reactions what would you expect to happen when an aqueous solution of barium hydroxide is mixed with an aqueous solution of a potassium salt?—*Aberdeen Univ.*

10. Give an account of the chief forms in which calcium carbonate is found. Ten grams of quicklime are slaked, mixed to a paste with water, and added to an excess of sodium carbonate. How much caustic soda is formed?—*Owens Coll.*

11. How is potassium chloride converted into (a) caustic potash, (b) potassium chlorate? Starting from *carrollite*, the double chloride of magnesium and potassium, show briefly how the metals magnesium and potassium can be prepared.—*Owens Coll.*

12. What is meant by (a) spark spectrum, (b) absorption spectrum, (c) reversed lines? How would you distinguish between the spectra of the metals of the alkaline earths?—*Cape Univ.*

CHAPTER XX

ELECTRICAL ENERGY

§ 1. Electrochemical Series of the Elements.

An electrochemical series is obtained by arranging substances in accord with their electrical properties, and this series is better than any other for giving a general idea of chemistry.—J. J. BERZELIUS (1825).
The metals precipitate one another after a certain order.—T. BERGMAN (1779)

NEAR the beginning of the 4th century Zosimus mentioned the fact that when iron is immersed in a solution of a copper salt, the iron acquires a coating of copper; and miners have frequently noticed that their iron tools become coated with copper when brought in contact with the water percolating through certain mines. We now know that such water may hold in solution copper sulphate from the oxidation of ores containing copper sulphide associated with iron sulphide, and we have also learned that when copper is deposited, an equivalent amount of iron passes into solution. The reaction is represented in symbols, $\text{Fe} + \text{CuSO}_4 = \text{Cu} + \text{FeSO}_4$. Curiously enough, even as late as the 16th century Paracelsus attributed the phenomenon to the transmutation of iron into copper. The transmutation hypothesis certainly appeared a very plausible explanation of the facts. As T. Bergman emphasized near the middle of the 18th century in his *De præcipitationis metallicis*:

The man who first saw a metal corroded by a limpid menstruum in such a manner that a body so extremely ponderous and so opaque should gradually and entirely disappear, and afterwards, upon the addition of a suitable precipitant to a liquid which appeared to be simple and homogeneous, saw the metal separate and again come into view, that man, I say, who first saw this, must have been struck with astonishment and admiration. Persons accustomed to these wonderful phenomena neglect, perhaps too much, the accurate investigation of them, though these operations are of the highest importance, and form as it were the whole of the effective part of chemistry.

Metallic magnesium will displace hydrogen from dilute acids: $\text{Mg} + \text{H}_2\text{SO}_4 = \text{MgSO}_4 + \text{H}_2$; or in the language of the ionic hypothesis: $\text{Mg} + 2\text{H}^+ + \text{SO}_4^{--} = \text{Mg}^{++} + \text{SO}_4^{--} + \text{H}_2$. Magnesium will also precipitate zinc from a solution of a zinc salt: $\text{Mg} + \text{ZnSO}_4 = \text{MgSO}_4 + \text{Zn}$; or in terms of the ionic hypothesis: $\text{Mg} + \text{Zn}^{++} + \text{SO}_4^{--} = \text{Mg}^{++} + \text{SO}_4^{--} + \text{Zn}$. Zinc in turn will precipitate iron from iron salts; iron will precipitate copper from copper salts; copper will precipitate silver from silver salts, etc. By treating the different metals in a similar manner, it has been found possible to arrange them in a series such that any metal in the list will usually displace most of those which follow it, and be displaced by most of those which precede it, thus:

$\text{Mg} \rightarrow \text{Al} \rightarrow \text{Mn} \rightarrow \text{Zn} \rightarrow \text{Cd} \rightarrow \text{Fe} \rightarrow \text{Co} \rightarrow \text{Ni} \rightarrow \text{Sn} \rightarrow \text{Pb} \rightarrow \text{Bi} \rightarrow \text{etc.}$

The order and behaviour of the metals vary a little with the nature and concentration of the solutions. Again, when zinc is treated with dilute acids under suitable conditions in a voltaic cell (Fig. 4), so as to eliminate disturbing effects, the reaction produces an electric current at a certain voltage. If the zinc be replaced by some metals—aluminium, magnesium, etc.—the voltage of the cell is increased; and conversely, if the zinc be replaced by other metals—cadmium, iron, cobalt, etc.—the voltage of the cell is diminished. It is thus possible to arrange the elements in a series representing the potential difference in volts which is developed between the metals and solutions of their salts.

The order here is virtually the same as the above list showing the order in which the elements displace one another from their salts. The list of the elements so arranged is called the electrochemical series. A more complete list is indicated in Table XXIV.

The order may vary a little with different solutions, or at different temperatures. Thus, tin precipitates lead from its solution in acetic acid, and lead precipitates tin from its solution in nitric acid. Calcium iodide at about 740° is reduced by sodium, but over 800° sodium iodide is reduced by calcium. Secondary reactions may prevent the precipitation of the metal. In many cases, the displacement is so complete that the reaction is employed in quantitative analysis. The further apart two metals in the series: (i) the greater the electromotive force of the current generated when the two elements are used as plates in a voltaic cell; (ii) the greater the amount of heat liberated when the displacement occurs, *e.g.* when zinc precipitates silver more heat is evolved than when it precipitates tin; (iii) similar remarks apply, *mutatis mutandis*, to the speed of precipitation; (iv) the greater the amount of heat or electrical energy required for the decomposition of their compounds; (v) the greater their chemical affinity for one another; etc.

A similar table would be obtained if the elements were arranged in the order of their chemical activity. Thus, (1) The earlier members on the list oxidize or rust on exposure to the air. (2) The oxides of the metals

TABLE XXIV.—ELECTRO-CHEMICAL SERIES OF THE ELEMENTS.

Cæsium
Rubidium
Potassium
Sodium
Lithium
Barium
Strontium
Calcium
Magnesium
Aluminium
Chromium
Manganese
Zinc
Cadmium
Iron
Cobalt
Nickel
Tin
Lead
Hydrogen
Antimony
Bismuth
Arsenic
Copper
Mercury
Silver
Palladium
Platinum
Gold
Iridium
Rhodium
Osmium

"Metals."

Silicon
Carbon
Boron
Nitrogen
Selenium
Phosphorus
Sulphur
Iodine
Bromine
Chlorine
Oxygen
Fluorine

"Non-metals."

succeeding manganese are reduced to metals when heated in a stream of hydrogen, while the metals which precede manganese under the same conditions, may be reduced to lower oxides, but not to the metallic condition. (3) The oxides of the metals mercury to osmium may be decomposed into their elements by simply heating them to a comparatively low temperature. (4) The metals preceding hydrogen on the list can give hydrogen when treated with acids, although secondary actions may simultaneously lead to the formation of some product other than hydrogen. The metals succeeding hydrogen do not usually displace hydrogen from the acids. (5) With the possible exception of tin and lead (metals close to hydrogen) the free elements are rarely, if ever, found in nature excepting possibly in meteorites. This arises from the fact that natural waters containing carbonic and other acids in solution attack these metals; consequently, even if these elements were produced by subterranean agents—volcanic or otherwise—they must succumb to attack by natural waters.

It will be noticed that the series only refers to the action of the *free elements*, and it has no direct reference to the mutual action of *chemical compounds* of the elements upon one another. The order of the elements in the electro-chemical series depends to some extent upon the temperature as well as on the nature and concentration of the electrolyte. For example, zinc and copper behave in what appears to be an abnormal manner in the presence of potassium cyanide. Thus copper and iron will precipitate zinc from potassium zinc cyanide, whereas zinc will precipitate copper from copper sulphate; and iron from neutral ferrous sulphate. Again, silver will displace hydrogen from aqueous hydriodic acid; copper will precipitate nickel from sodium nickel chloride; and platinum will liberate hydrogen from aqueous solutions of potassium cyanide.

Electro-affinity.—The idea has been expressed another way. It is assumed that the ions hold their charges with different degrees of tenacity. The ions— K^+ , Na^+ , NO_3^- , Cl^- , etc.—which hold their charges very tenaciously, are called strong ions: and ions— Hg^{++} , Ag^+ , OH^- , Cy^- , etc.—which readily lose their charge, are called weak ions. The degree of tenacity with which the ions of an element hold their charges has been called the electro-affinity of the element. Ions with strong electro-affinity are difficult to prepare in a free state, and conversely. If an element with a strong electro-affinity comes in contact with the ion of an element with a weak electro-affinity, the charge on the latter passes over to the former. Thus, zinc has a stronger electro-affinity than copper, and, in consequence, as indicated above, zinc will precipitate copper from solutions of its salts: $Zn + Cu^{++} = Zn^{++} + Cu$. Zinc also has a stronger electro-affinity than hydrogen, and consequently zinc dissolves in dilute acids with the evolution of hydrogen: $Zn + 2H^+ = Zn^{++} + H_2$. Similarly, chlorine has a stronger electro-affinity than bromine, and bromine a stronger electro-affinity than iodine. In consequence, chlorine will displace bromine from aqueous solutions of the bromides: $Cl_2 + 2K^+ + 2Br^- \rightleftharpoons 2K^+ + 2Cl^- + Br_2$; and bromine will displace iodine from the iodides: $Br_2 + 2K^+ + 2I^- \rightleftharpoons 2K^+ + 2Br^- + I_2$.

§ 2. Solution Pressure—Contact Differences of Potential.

Whenever there is free energy present, or whenever there is potential energy seeking to become free (as where chemical affinities can come into play) in the presence of differential molecular structures or conditions, there energy takes the form of electrical potential.—J. T. SPRAGUE.

Every metal, except, of course, the last in the series, will displace those that succeed it in the electrochemical series, and it is inferred as an hypothesis that each of these metals has a tendency to become ionic. This tendency is hypothetically regarded as a solution pressure which drives the ions of the metals into solution. This pressure must be greatest with the metals at the caesium end, and least with the metals at the osmium end of the series. Conversely, the tendency of positive metal ions in solution to reprecipitate on the negative electrode must be least at the caesium end of the series, and greatest at the osmium end. The ionic hypothesis assumes that this back or deposition pressure represents the osmotic pressure of the ions.

The dissolution or ionization of a metal has been compared with the tendency of different liquids to vaporize at any given temperature. Just as a liquid in a closed vessel will evaporate until the number of molecules leaving the surface of the liquid in a given time is equal to the number of molecules returning to the liquid, so W. Nernst (1889) has suggested that a metal when placed in contact with water, or any other solvent, sends positively charged ions into the solvent, and itself acquires a negative charge. The ionization of the metal, so to speak, is supposed to continue until the concentration of the metallic ions in the liquid has attained a certain value, when a state of equilibrium ensues. The number of ions passing into the solution is then equal to the number reprecipitated on the surface of the metal. Direct proof of the presence of iron ions in purified water which has been in contact with the highly purified iron is wanting. The evidence is indirect, or rather hypothetical.

When zinc is immersed in dilute hydrochloric acid, the H^+ ions which come in contact with the zinc plate lose their charge, and positively charged zinc ions pass into solution. If a stick of metallic zinc be dipped in a saturated solution of zinc sulphate, the solution and osmotic pressures are balanced, and no action occurs; but if a stick of metallic zinc be placed in a dilute, say normal, solution of zinc sulphate, the solution pressure is greater than the osmotic pressure, and positively charged zinc ions pass from the zinc rod into the solution. In consequence, the zinc acquires a negative charge, and the solution a positive charge, in agreement with the fact that zinc usually acquires a negative charge when immersed in a solution of its own salt. Similar remarks apply to aluminium, iron, etc. Conversely, if the solution pressure be less than the osmotic pressure of the ions, as appears to be the case with a stick of metallic copper immersed in a solution of copper sulphate, copper ions will be deposited on the metal, and the solution will acquire a negative charge while the metal acquires a positive charge. This also appears to be the case with the metals silver, mercury, etc.

The electrical effect, or the contact difference of potential, produced when the different metals are immersed in a normal solution of their sulphates has been measured. A few of the results are here indicated :

	Volts.
Magnesium	+1.214
Zinc	+0.493
Cadmium	+0.141
Hydrogen	0
Copper	-0.606
Silver	-1.048

The "+ 0.493" opposite zinc means that if metallic zinc be immersed in a normal solution of zinc sulphate, the solution will acquire a positive charge, and the metal a negative charge; and the difference of potential between the solution and the metal will be 0.493 volt. With metallic copper and a solution of copper sulphate, the solution will be charged negatively, and the copper positively such that the difference of potential between the solution and the metal will be 0.606 volt.

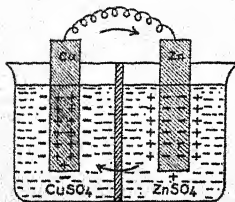


FIG. 152.—Daniell's Cell
(Diagrammatic).

If a normal solution of copper sulphate be separated by a porous partition, Fig. 152, from a normal solution of zinc sulphate, and if a rod of copper immersed in the copper sulphate be connected by a wire with a rod of zinc immersed in the zinc sulphate (Fig. 152), the zinc pole on the right of the diagram acquires a negative charge, *ex hypothesi*, on account of the departure of positively charged ions from its surface, and the copper pole on the left acquires a positive charge on account of the departure of negatively charged copper ions from its surface. In consequence, an electrical current will flow through the connecting wire from the positively to the negatively charged pole and pass in the converse direction through the liquid. This action continues until all the zinc is dissolved or all the copper precipitated. The relative solution pressures of the two metals decide the magnitude of the resultant electromotive force of the current, and this is the difference of the two effects. The resultant electromotive force for the zinc : copper couple just described is $+0.493 - (-0.606) = 0.493 + 0.606 = 1.099$ volts. The combination just described represents the so-called Daniell's cell (1836). In reality, the Daniell's cell contains the zinc rod with the sulphuric acid or zinc sulphate solution in a porous pot, and the copper plate with the copper sulphate solution in the surrounding jar, as illustrated by the drawing of an uncharged cell in Fig. 153. The reactions are symbolized molecularly : $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$; and instead of the hydrogen atoms being liberated on the negative copper plate, they are exchanged for copper atoms in the cupric sulphate solution : $\text{H}_2 + \text{CuSO}_4 = \text{H}_2\text{SO}_4 + \text{Cu}$, and the copper is deposited on the outer copper plate. The round of changes results in the growth of the copper plate, and in the wearing away of the zinc plate; zinc sulphate increases, copper sulphate decreases. When a zinc sulphate solution is used instead of sulphuric acid, the copper is exchanged chemically for the zinc : $\text{Zn} + \text{CuSO}_4 = \text{Cu} + \text{ZnSO}_4$. There are many other modifications of

Daniell's cell; and numerous other types of cell with different "poles" and different solutions.

The *quantity* of electricity (coulombs) produced by Daniell's cell depends upon the amount of zinc consumed (Faraday's law); and the *rate* at which electricity is developed (ampères) depends upon the rate at which the zinc is consumed in the cell.

The *difference of potential*, or the electrical pressure, cannot exceed 1.099 volts for the given solutions. The product of the number of coulombs into the number of volts gives the amount of electrical energy produced in a given time. This product expresses the electrical energy in joules: joules = coulombs \times volts.

If the term *ampère* be employed to represent a current equivalent to one coulomb per second, the product of the number of ampères into the number of volts gives the amount of electrical energy produced by the cell. This product expresses the rate of production of electricity by the cell in watts. That is watts = amps. \times volts. A joule of electrical energy per second represents one watt.

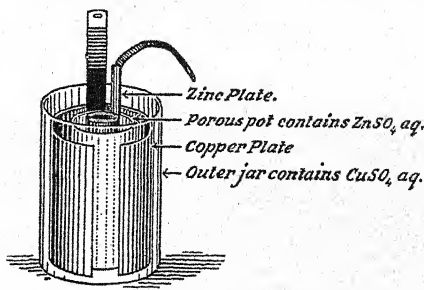


FIG. 153.—Daniell's Cell.

EXAMPLES.—(1) The electrolysis of 36.5 grams of hydrochloric acid requires 96,540 coulombs of electricity at 1.31 volts. Hence the electrical energy needed for this work is $96,540 \times 1.31 = 126,567$ units, or, defining a joule as the unit of electrical energy consumed by a current of one coulomb working at a pressure of 1 volt; here, joules = volts \times coulombs. Hence, the electric energy needed to decompose 36.5 grams of hydrochloric acid is 126,567 joules. Since one watt is equivalent to a joule per second, the answer will be 126,567 watts (second).

(2) A current of 100 volts and 1.5 amps. passes through a system; find the energy consumed per second in kilowatts. Here $100 \times 1.5 = 150$ watts or 0.15 kilowatts (second).

(3) If 100 watts be expended by a current of 5 amps., what is the pressure required to drive the current through the circuit? Ansr. 20 volts.

(4) How many grams of silver will be deposited from a suitable electrolyte by a current of 100 amps. running for one hour? Ansr. 303.6 grams.

Although the difference of potential of a given cell, say, a Daniell's cell, is not affected by variations in the size or shape of the poles, or upon the quantity of liquid in the cells, the difference of potential is altered by changing the concentration of the solutions. In general, the difference of potential between a metal and a solution of one of its salts is greater with increasing dilution. A tenth normal solution of zinc sulphate, for instance, was found to give a difference of potential of 0.551 volt, whereas with a normal solution a potential difference of 0.493 volt is obtained as indicated above.

If two rods of zinc be separately placed in a solution of zinc sulphate, the difference of potential in both "tends" to drive an electric current from the metal to the solution with a pressure of 0.493 volt. If both rods be joined by a wire, no electric current will flow because the two equal

forces are oppositely directed. On the other hand, if the zinc rods be dipped in solutions of a *different* concentration, the two contact differences of potential will be different, and an electric current will flow from the concentrated solution to the dilute solution outside the cell as indicated in Fig. 154. Here a normal solution of zinc chloride is supposed to be placed in one vessel, *A*, and a decinormal solution of zinc chloride in the other vessel *B*. Zinc rods connected by a copper wire and galvanometer are dipped into the solutions, as illustrated in the diagram, and the two cells are connected by a syphon tube *S*. The difference of potential of the zinc in the normal solution is $+0.493$, and in the more dilute solution $+0.551$ volt. Hence an electric current tends to pass from the metal to the dilute solution with a force of $+0.551$ volt, and from the metal to the concentrated solution with a force of 0.493 . The resultant pressure is therefore $0.551 - 0.493 = 0.058$ volt, and this represents the electromotive force of the combination. Cells in which the electromotive force is generated by the difference potential of two plates immersed in solutions

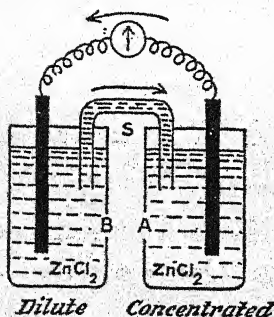


Fig. 154.—Concentration Cell.

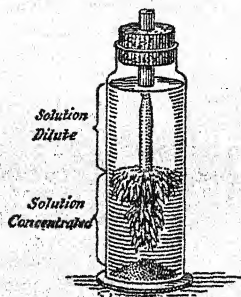


Fig. 155.—Experiment on Concentration Cells.

of the same salt at different concentrations are called **concentration cells**. The chemical action which occurs in the two cells tends to bring the two solutions to the same concentration. The action is made clear by the experiment illustrated in Fig. 155. A layer of a concentrated solution of stannous chloride in hydrochloric acid, about 10 cm. deep, is placed at the bottom of a cylinder, and above this a layer of a dilute solution. A rod of metallic tin is fixed through a hole in the cork so that it is suspended axially in the liquid in the cylinder. The rod of tin thus represents both electrodes and connecting wire of a concentration cell. Tin is dissolved by the more dilute solution, and precipitated from the more concentrated solution. The diagram illustrates the appearance of the rod of tin after the vessel has stood a couple of days. Cadmium and cadmium chloride could be used; zinc and zinc chloride do not work so well; nickel gives a negative result, probably because the surface becomes polarized; antimony and bismuth also give a negative result, probably because they are not attacked by dilute hydrochloric acid.

There is another interesting feature about a concentration cell. If an external electromotive force be applied so as to force an electric current to

pass in a reverse direction to that which the combination normally furnishes when it is employed as a voltaic cell, the chemical actions will be reversed, and the difference in the concentration of the two solutions will be augmented. Such combinations are called **reversible cells** in contradistinction to **irreversible cells** in which the original condition cannot be restored by sending a current through the cell in a reverse direction to the current normally delivered by the cell. The cell illustrated in Fig. 4 is an irreversible cell; Daniell's cell, Fig. 153, and the concentration cell, Fig. 154, are reversible cells.

Couples.—If metallic zinc dissolving in, say, dilute sulphuric acid, be in contact with a piece of copper or platinum the rate of dissolution of the zinc is augmented. The combination Zn : Pt is called a *couple*; it is really a small galvanic cell with zinc and, say, platinum electrodes connected together by metallic contact. Much of the hydrogen is evolved from the surface of the platinum as might be expected from the description of Fig. 4.

Sacrificial metals.—We have seen that any metal in the electrochemical series can be made one plate of a cell against a metal lower down in the series. Zinc, for instance, can be made the positive plate against a negative plate of iron, tin, lead, etc.; and iron the positive plate against a negative plate of tin, lead, etc. The further apart the elements in the series, the greater the electromotive force of the combination. *Tin-plate* is iron or steel coated with a thin layer of tin. If a little moisture be precipitated on the surface in contact with both the iron and the tin, the moisture, with its dissolved carbonic acid, dissolves the iron, producing salts of iron; these ultimately form rust (*q.v.*). The iron is covered with a layer of tin to protect it from rust, but if there be a flaw in the protecting surface of tin so as to expose the underlying iron, rusting takes place more rapidly than if the iron had not been tinned at all. The tin remains untarnished. Zinc is also used as a protecting layer over the surface of thin iron plates—*galvanized iron*. The voltaic action developed when the protecting layer is damaged is much less than when tin is used. These facts can be illustrated by fitting up a cell like Fig. 4 with iron and tin plates, and another cell with iron and zinc plates. Water saturated with carbon dioxide is used in both cells.

A feeble electric current will flow from the tin to the iron outside the cell in one case, and from the iron to the zinc in the other as illustrated graphically in the adjoining diagram. In the iron : tin cell, iron dissolves and rusting occurs; while in the iron : zinc cell, the zinc dissolves and no rusting occurs as long as the circuit is closed. These results might almost have been predicted from our study of Table XXIV., p. 433. An iron : lead cell behaves like an iron : tin cell. Iron railings are often fixed in a bed of lead, the iron corrodes first and the lead remains intact. H. Davy (1824) once proposed to prevent the corrosion of the copper sheathing of ships by fixing pieces of metallic zinc here and there on the sheathing. The zinc was corroded and the copper preserved.¹ In all these cases it has been fancifully said that one metal



¹ The copper then ceased to poison the barnacles, and the bottom fouled as if the wood had not been sheeted with copper.

is sacrificed to ensure the safety of the other; and all the cases quoted are examples of galvanic couples: Fe : Sn; Zn : Fe; Fe : Pb; and Zn : Cu.

§ 3. The Ionic Hypothesis in Difficulties.

The knowledge of nature as it is—not as we imagine it to be—constitutes true science.—PARACELSUS.

There are some enthusiasts who claim that "all chemical reactions are reactions between ions; molecules as such do not react at all." This statement is not quite in harmony with known facts. The same might be said of the assumption that "chemical activity is proportional to the number of available ions." L. Kahlenberg (1902 *seq.*) has brought forward so large a number of exceptions to these statements that it will be necessary to modify the hypothesis very materially before it can be accepted as an accurate description of the facts. Some chemical reactions proceed very rapidly in solutions which are considered to be non-conductors of electricity, and which, *ex hypothesi*, are almost free from ions. For instance, dry hydrogen chloride precipitates chlorides from benzene solutions of the oleates of copper, cobalt, and nickel; dry hydrogen sulphide precipitates sulphides from benzene solutions of the same salts and of arsenic chloride. All this in spite of the fact that these solutions do not conduct electricity appreciably. Again, dry ammonia does not unite with dry hydrogen chloride, but union does take place if a trace of non-conducting benzene vapour be present. One metal can displace another from a non-conducting solution in a non-aqueous medium. Thus metallic lead, zinc, tin, silver, iron, etc., will precipitate metallic copper from solutions of various salts in carbon disulphide, carbon tetrachloride, ether, alcohol, etc. Hence, in spite of the ionic hypothesis, chemical reactions do take place in non-conducting solutions, and these reactions are similar in result and speed to those which occur in conducting aqueous solutions. The ionic hypothesis cannot, therefore, ignore these observations if it is to win a permanent place among the conquests of science. There are ominous signs that the ionic hypothesis will have to go. As O. D. Chwolson emphasized, there is as yet no solidly established theory of solutions which will take account of all the known phenomena, and on which reliance can be placed with safety.

§ 4. Polarization—Back Electromotive Force.

The source of the chemical energy in a galvanic cell is certainly the chemical action, a correction being applied for any reversible heat which the cell absorbs from or gives up to its surroundings.—W. C. D. WHETHAM.

Just as an ordinary steam boiler is a device for transforming the chemical energy of burning coal into mechanical energy, so can a voltaic cell be regarded as a device for converting chemical energy into electrical energy. In one of the simplest cases, when the simple cell Zn : H₂SO₄ : Pt, of Fig. 4, is working, hydrogen is evolved, mainly from the surface of the platinum. The chemical action is vigorous at first, but gradually diminishes in intensity, and finally nearly stops altogether. The curve, Fig. 154, shows the electromotive force of such a cell working with a

resistance of about ten ohms in the external circuit after different intervals of time. The rapid drop from an initial electromotive force of 1.3 volts to about half a volt in five minutes is indicated by the rapid descent of the curve. After five minutes, the electromotive force remained fairly constant at about 0.4 volt, that is nearly 66 per cent. below the initial value. The effect is easily illustrated by connecting an electric bell with such a cell. The bell rings loudly at first, but gradually weakens, and finally stops. If the platinum plate be then removed, the surface will be found covered with a layer of bubbles of hydrogen gas, which has remained on the surface of the plate instead of passing away. If the circuit is broken, the bubbles of gas gradually dissipate from the platinum plate, and the cell then resumes its former electromotive force when the circuit is closed. This temporary reduction in the electromotive force of a cell is said to be due to the polarization of the cell. Polarization may be regarded as

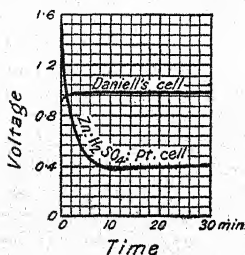


Fig. 156.—Voltage Drop of Voltaic Cells.

electro-chemical fatigue which causes a modification of one or both of the plates, or by the exhaustion of the solution produced during the working of the cell.

In consequence of this phenomenon, commercial cells have some provision for depolarization, that is, for preventing the accumulation of gas on the negative plate. In Daniell's cell, the variation in the electromotive force of the working cell is chiefly due to changes in the concentration of the solution surrounding the battery plates. The electromotive force is therefore nearly constant. This is illustrated by the curve

shown in Fig. 156, where a Daniell cell was allowed to work for half an hour against a resistance of 10 ohms in the external circuit. A comparison of this curve with that of the Zn : H₂SO₄ : Pt cell emphasizes the constancy of the current delivered by the Daniell cell.

W. Grove's cell has a porous pot like Daniell's cell, but the pot contains a platinum plate and nitric acid; zinc and sulphuric acid are contained in the outer compartment. The hydrogen produced by the action of the sulphuric acid on the zinc instead of accumulating on the platinum plate is oxidized by the nitric acid and the latter is reduced—say, $\text{H}_2 + 2\text{HNO}_3 = 2\text{H}_2\text{O} + 2\text{NO}_2$. Owing to the high cost of platinum, J. T. Cooper proposed to replace it by carbon, and this combination is known as *R. Bunsen's cell*. In *A. Smee's cell*, the platinum plate is replaced by a silver plate covered with finely divided platinum. The idea is to make the hydrogen bubbles collect on the roughened surface from which they escape more freely than from a plain surface. The voltage drop with this cell is large. In *G. Leclanche's cell* a zinc rod and carbon plate are clamped between two blocks of compressed manganese dioxide and granulated carbon. This combination is dipped in a solution of ammonium chloride. When the circuit is closed, the ammonium chloride attacks the zinc: $\text{Zn} + 2\text{NH}_4\text{Cl} = \text{ZnCl}_2 + 2\text{NH}_3 + \text{H}_2$; the hydrogen is burnt by the manganese dioxide: $\text{H}_2 + 2\text{MnO}_2 = \text{Mn}_2\text{O}_3 + \text{H}_2\text{O}$. The cell does not give a constant current except for a short time, but if left standing it rapidly recovers. Millions of these cells are used where a current is required only for a short time—electric bells, etc. There are numerous modifications. In *R. Bunsen's dichromate cell*, zinc and carbon plates dip in a solution of sulphuric acid mixed with chromic acid, or an alkali dichromate. The zinc is attacked by the liquid when the cell is not in use so that there is an arrangement for lifting the plates out of the liquid when the cell is not in use. The object of the chromic acid is to burn the hydrogen and prevent its accumulating on the

carbon plate; the chromic acid is reduced to chromic oxide. The reaction at the zinc plate is therefore $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$, and then $3\text{H}_2 + 2\text{CrO}_3 = \text{Cr}_2\text{O}_3 + 3\text{H}_2\text{O}$. The solution contains sulphuric acid so that the chromic oxide forms chromic sulphate. In *L. Clarke's standard cell*, the positive plate of purified zinc is immersed in a saturated solution of zinc sulphate, and the negative plate is mercury in contact with mercurous chloride. The voltage is remarkably constant at about 1.434 volts under standard conditions. There are many modifications; in *E. Weston's cell*, for instance, there is metallic cadmium and cadmium sulphate in place of zinc and zinc sulphate.

Again, if a current exceeding two volts be directed through an electrolytic cell (Fig. 6) containing dilute sulphuric acid, and fitted with two platinum plates and a galvanometer in circuit, bubbles of gas are disengaged at the two electrodes, oxygen at the anode, hydrogen at the cathode. The direction of the current is indicated by the deflection of the needle of the galvanometer. Now let the battery be cut out of the circuit, and the electrodes be joined directly with the galvanometer. The deflection of the needle shows that a feeble current passes in an *opposite* direction to that which occurred when the battery was in circuit. An examination of the plates of the polarized electrolytic cell shows that gaseous films are present. Obviously, therefore, after a current has passed through such a cell for a short time, the plates—originally quite similar—are no longer alike. The plates are polarized with different gases. The gases adhere to the surface and penetrate the interior of the plates. The plates then behave as if they were made of two different materials. Contact differences of potential are established. We have in fact a voltaic cell, $\text{O}_2 : \text{H}_2\text{SO}_4 : \text{H}_2$, which furnishes a current flowing in an opposite direction to the original current. The cell acts as a kind of "accumulator" of electrical energy until the gases absorbed by the plates are used up. This does not take long. The polarization of the plates of an electrolytic cell thus makes them behave like two different metals which exert a *back electromotive force* opposing the electromotive force of the battery.

§ 5. Decomposition Voltages.

In a mixture of electrolytes, those substances are set free which absorb in becoming free the least intrinsic energy or lowest voltage.—J. T. SPRAGUE.

Suppose a current of $\frac{1}{2}$ volt be sent through the electrolytic cell, Fig. 6, containing normal sulphuric acid and fitted with platinum plates in circuit with a galvanometer. The current passes through the cell for an instant as indicated by the "throw" of the galvanometer needle, and then the quick drop to nearly zero. The hydrogen and oxygen developed on the plates sets up a back electromotive force of nearly $\frac{1}{2}$ volt which very nearly stops the current. A minute steady current—residual current—does flow through the system, but this is only just sufficient to maintain the polarization, since if no current at all passes through, the plates would gradually depolarize owing to the dissipation of the gases from the plates.

If the current be now raised to 1 volt, a similar state of things prevails. The amount of oxygen and hydrogen adhering to the plates increases; and the increased polarization raises the back electromotive force to very nearly one volt. The residual current passing through the cell is slightly larger than before. This is required to maintain the polarization.

If the current is now raised to 1.7 volts, the electrodes become saturated with hydrogen and oxygen gases. Polarization reaches a maximum value, and the back electromotive force also attains its maximum value. Hence any further increase in the applied electromotive force is available for electrolysis, 1.7 volts are the minimum needed for steady electrolysis. If 2 volts are passed through the system, there is a back electromotive force of about 1.7 volts, and the "excess" or "residual" current, 0.3 volt, is the effective electromotive force available for the production of current, and the steady evolution of gases from the electrodes.

The facts here described can be exhibited very concisely by plotting the applied electromotive forces as ordinates and quantities of electricity passing through the system as abscissae. Fig. 157 shows the results with normal solutions of sulphuric acid, hydrochloric acids, and silver nitrate. The "residual" current flowing through the cell with normal sulphuric acid rises very slowly with increasing voltages until the driving force reaches 1.67 volts. There is then a sudden change in the direction of the curve. Increasing electromotive forces now augment the quantity of electricity passing through the system, and also the amount of electrolysis. Normal hydrochloric acid gives a similar break at 1.31 volts; and silver nitrate, one at 0.70 volt.

The minimum electromotive force required to cause steady electrolysis in any solution is called the decomposition voltage or discharge potential. The decomposition voltages for a few acids, bases, and salts are shown in Table XXV.

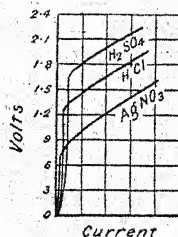


FIG. 157.—Effect of an Increasing E.M.F. on some Electrolytes.

TABLE XXV.—DISCHARGE POTENTIALS OF SOME ELECTROLYTES.

Electrolytes. (Normal solutions.)	Decomposition voltages. (Volts.)
Salts : Zinc sulphate	2.35
Nickel sulphate	2.09
Lead nitrate	1.52
Silver nitrate	0.70
Acids : Sulphuric acid	1.69
Hydrochloric acid	1.31
Nitric acid	1.69
Phosphoric acid	1.70
Bases : Sodium hydroxide	1.69
Potassium hydroxide	1.67
Ammonium hydroxide	1.74

While the values for the metallic salts vary from metal to metal, the acids and bases have a decomposition voltage approaching 1.7 volts, and the products of the electrolysis are oxygen and hydrogen. Those acids

which have a lower decomposition voltage usually give off other products on electrolysis, and attain the final value—1.7 volts—on further dilution. Thus hydrogen and chlorine are evolved when the strength of the hydrochloric acid exceeds 2*N*-HCl, and the decomposition voltage of the 2*N* acid is 1.26 volts. The voltage steadily rises with increasing dilution until, with $\frac{1}{32}$ *N*-HCl the decomposition voltage is 1.69, and hydrogen and oxygen are the products of electrolysis. Not only do the numbers vary with concentration, within certain limits, as exemplified in the case of hydrochloric acid, but also with the nature of the electrodes. The decomposition voltage of normal sulphuric acid, for example, with polished platinum electrodes is 1.67 volts, whereas with platinum electrodes covered with platinum black, the decomposition voltage is 1.07 volts.

The contact potential between metallic zinc and a normal solution of a zinc salt, — 0.493 volt, shows that when a zinc ion is deposited on a zinc electrode it conveys a positive charge to the electrode and so lessens the negative charge there present. The system is only in equilibrium when the zinc electrode is negatively charged to a potential of — 0.493 volt. If, therefore, zinc is to be deposited in an electrolytic cell, this difference of potential must be counterbalanced by the current. Hence contact differences of potential may also be regarded as decomposition voltages.

The discharge potentials of a few anions and cations are indicated in Table XXVI., which may be compared with Table XXV. The numbers refer to normal solutions. The prefix refers to the electrical state of the electrode in the presence of a normal solution of its ions, say, 32.5 grams of zinc per litre. Some of the numbers have not been measured directly. For instance, the number for zinc sulphate has been obtained by extrapolation, since, according to the conductivity measurements, only 23 per cent. of zinc sulphate is ionized in normal solutions.

TABLE XXVI.—DISCHARGE POTENTIALS OF SOME ANIONS AND CATIONS.

Cation.	Charge on metal volts.	Anions.	Charge in volts.
Zn ⁺⁺	— 0.493	I [']	+ 0.797
Fe ⁺⁺	— 0.063	Br [']	+ 1.270
Ni ⁺⁺	+ 0.049	O ^{''} (in acid)	+ 1.396
Sn ⁺⁺	+ 0.085	Cl [']	+ 1.694
Pb ⁺⁺	+ 0.129	OH ['] (in acid)	+ 1.96
H [']	+ 0.277	OH ['] (in bases)	+ 1.16
Cu ⁺⁺	+ 0.608	NO ₃ [']	+ 1.75
Hg ⁺⁺	+ 1.027	SO ₄ [']	+ 1.9
Ag ⁺⁺	+ 1.048	HSO ₄ [']	+ 2

Just as different electrical pressures (E.M.F.) are needed to produce in different solutions equivalent amounts of chemical change, so different chemical reactions in a voltaic cell generate different amounts of electrical energy, and produce currents with different electromotive forces. During electrolysis a difference of electrical pressure must be continuously supplied because the current is consumed, so to speak, by the separation of chemically equivalent quantities of matter (Faraday's law). In a voltaic cell

electrical energy is produced, so to speak, from the chemical energy of the dissolving zinc. The question whether or not a given supply of electrical energy can start electrolysis is determined by the intensity, pressure, or voltage of the current. The total supply of available electrical energy does not matter. Although a given quantity of electricity, say 96,540 coulombs, will separate chemically equivalent quantities of different electrolytes, these 96,540 coulombs must be supplied at definite pressures before electrolysis can take place. In other words, just as different compounds decompose at different temperatures and this quite independent of the total quantity of available heat, so electrical energy at different voltages is needed for the decomposition of different electrolytes.

The total amount of electrical energy required for the liberation of chemically equivalent quantities of different electrolytes can be approximately determined by multiplying the 96,540 coulombs (or one farad) of electricity by the voltage needed for electrolysis. Hence the decomposition voltage is a measure of the energy needed for the decomposition of a gram equivalent of a given electrolyte, and the product of the quantity of electricity into its electromotive force not only represents the energy of a battery, but it also measures the chemical energy which is transformed by the battery into electrical energy. As indicated above, a joule, the unit of electrical energy, is numerically equivalent to the product of one volt into one coulomb. Further, a joule is equivalent to 0.2382 calorie of thermal energy. Hence, just as thermochemistry would write $\text{Na} + \text{Cl} = \text{NaCl} + 97,900$ calories, so electrochemistry writes $\text{Na} + \text{Cl} = \text{NaCl} + 411,000$ joules. The amount of heat evolved or absorbed when a given compound is decomposed can be measured; the minimum amount of electrical energy required to decompose a given compound must be at least equivalent to the amount of heat developed when the necessary constituents unite to form the original compound. As a first approximation, it may be assumed that the heat of formation of any given compound is a measure of the thermal equivalent of the electrical energy required to break up the compound by electrolysis: chemical energy expended \rightleftharpoons electrical energy produced. Otherwise expressed, thermal value of reaction \rightleftharpoons electrical energy produced. Both forms of energy are sometimes expressed in terms of the joule using the conversion factor: 1 joule = 0.24 calorie.

EXAMPLES.—(1) The heat of formation of sodium chloride is 97,900 calories: what is the equivalent electrical energy needed for the electrolysis of a gram equivalent of the fused salt, and what is the decomposition voltage required? Here, 97,900 calories are equivalent to $97,900 \div 0.24 = 411,000$ joules. But 96,540 coulombs will liberate chemically equivalent quantities of sodium and chlorine, and 411,000 joules are needed for this purpose. Consequently, since electrical energy = volts \times coulombs; $411,000 = \text{volts} \times 96,540$; or volts = 4.3. This means that in order to liberate 23 grams of sodium and 35.5 grams of chlorine from 58.5 grams of fused sodium chloride, 411,000 joules of electrical energy must be supplied at a minimum voltage electromotive force of 4.3 volts. The minimum voltages so calculated are usually a little higher than are needed in practice. It will be observed that the arithmetic may be summarized by the formula: volts = $0.043 \times \text{Cals.}$, where Cals. is employed to denote the heat developed by the reaction with an equivalent weight of the given compounds expressed in grams. *E.g.* the equivalent weight of water, H_2O , is half the molecular weight; the equivalent weight of aluminium chloride, AlCl_3 , is one-third the molecular weight.

(2) What number of volts are required for the decomposition of water? Given

34.2 Cals. are needed to separate water into a gram of hydrogen and 8 grams of oxygen, $0.043 \times 34.2 = 1.47$ volts.

(3) What voltage is required for the electrolysis of fused magnesium chloride, MgCl_2 , when the heat of formation is 217.3 Cals. ? Ansr. 4.7 volts.

(4) The heat of formation of calcium chloride is 169,800 calories. What electrical energy is required to decompose a gram-molecule, and what voltage will be needed ? Ansr. 707.5 kilowatts (seconds), 3.65 volts.

§ 6. Fractional Electrolysis—Magnus' Rule.

Among a number of mixed electrolytes there is a selective power at the electrodes which is based solely on the ratios of the voltages required to free the several ions.—J. T. SPRAGUE.

When a solution containing salts of different metals is subjected to electrolysis, there is a certain voltage at which one and only one of the metals will be deposited on the cathode—Magnus' rule (1856). If a mixed solution of nickel and copper sulphates, for example, be subjected to electrolysis, copper alone is precipitated when the applied electromotive force has reached 1.29 volts: the nickel is not precipitated, since its decomposition voltage is 1.95 volts. On the other hand, if a mixture of nickel and iron sulphates be similarly treated, a mixture of iron and nickel will be simultaneously deposited. The decomposition voltages of these salts are too close to allow an effective separation of the two elements by electrolysis. Hydrogen is also evolved during the electrolysis of these salts. This arises from the fact that the decomposition voltage of sulphuric acid—1.67 volts—renders it also susceptible to the influence of the same current as liberates nickel and iron.

Many useful methods of analysis are based upon these principles. In metallurgy too, electrolytic processes for refining metals—nickel, copper, lead, tin, silver, gold, etc.—have been developed. For example, in copper refining, as we shall soon see, anodes made of crude copper are dipped in a solution of copper sulphate acidified with sulphuric acid; the cathodes are sheets of pure copper. Zinc, iron, and copper from the anode pass into solution during electrolysis. The decomposition voltage is kept below that needed for the deposition of zinc and iron. In consequence, refined copper is deposited on the cathode. Other impurities affecting the crude copper are but slightly soluble in the electrolyte, and are deposited about the anode as a thin mud—"anode mud."

The decomposition voltage of an electrolyte is greater the more dilute the solution. The concentration of any given salt about the electrode naturally decreases during the process of electrolysis. Hence also the decomposition voltage for that particular salt in the mixed electrolyte also increases. When the concentration of the copper sulphate in a mixture of copper and nickel sulphates has become so small that the decomposition voltage of the dilute solution approaches that of nickel, any further electrolysis will bring down a mixture of both metals. There is, therefore, a limit to the process of electrolytic separation, just as there is a limit to the separation of substances in ordinary analysis. The limit in the former case is determined by the decomposition voltages of the respective metals; and in the latter case the limit is determined by the solubility of the precipitates in the given menstruum. The limiting concentration can be approximately determined from the rule: A decrease

of $\frac{1}{10}$ in the concentration of the electrolyte raises the decomposition voltage of any given ion $0.058/m$ volt, where m is the valency of the particular ion.

As the cation is deposited about the cathode, the loss in concentration is made up by diffusion from the surrounding electrolyte. To hasten diffusion, and prevent the undue attenuation of the electrolyte in the vicinity of the cathode, stirring by rotating one of the electrodes is sometimes used.

If the cathode be small in comparison with the anode, the solution about the former will be very much more quickly exhausted than if a larger cathode had been used. The decomposition voltage of the substance will rise in a proportional manner. Hence the larger the cathode the lower the "average" electromotive force needed for the deposition of the pure metal. It is convenient to call the quantity of electricity flowing through unit surface area, the current density at the electrode, in other words, "the number of amperes per unit surface." "Unit surface" is usually taken in the laboratory to be one square decimetre. The symbol $N.D._{100} = \frac{1}{2}$ means that a current of 0.5 amp. flows for every 100 sq. cms. of electrode surface.

EXAMPLE.—What was the current density at each electrode of an electrolytic cell when 4 sq. cms. of each electrode were immersed in the electrolyte and a current of 4.25 amperes was passed through the system for one hour? One square decimetre = 100 sq. cms. Hence, 1.0625 amps. passed per sq. cm.; or 106.25 amps. per sq. decimetre. The current density at each electrode was therefore 106.25 amps., or $N.D._{100} = 106.25$.

Current density is one of the most important factors in electrolysis, since it determines the character and nature of the products obtained at the different electrodes. Thus, by using a large current density and a concentrated solution of sulphuric acid, hydrogen, oxygen, ozone, and free sulphur can be obtained, whereas under ordinary laboratory conditions the latter substance does not appear.

§ 7. The Factors of Energy.

The term electromotive force in electricity is equivalent to the term chemical activity or affinity, just as the term quantity of electricity corresponds to the chemical notion of valency.—M. SALET (1867).

All the differences discoverable in the effects of electricity (obtained from different sources) may be owing to its being less intense but produced in larger quantity (from some sources rather than from others).—W. H. WOLLASTON (1801).

The idea developed in the preceding sections, that electrical energy is dependent upon two distinct factors, may now be extended further. Water will only flow from one vessel to another when there is a difference in the level of the liquid in the two vessels. The actual volume of the water in either vessel does not matter. Again, heat will only pass from one body to another when the temperature of the one is higher than the temperature of the other. The flow of heat is not determined by the quantity of heat in either the hot or the cold body. The heat in the furnace of a locomotive can do work, not because it is hot, but because it is hotter than its surroundings. If two reservoirs of gas be connected by a cylinder fitted with a sliding piston, the motion of the piston will not be determined by

the volume of the reservoir, nor by the quantity of energy contained in the gas, but it will be determined by difference in the pressure of the gas in the two cylinders. In this sense it can be said that every form of energy has a dual nature, for all the different forms of energy are of two dimensions in that they are compounded of two factors—mass of water and difference of level; thermal capacity and temperature; volume and pressure of gas. The one factor is called the quantity or capacity factor, and the other the intensity factor or strength.

Available energy = capacity (quantity) factor \times intensity (strength) factor.

Hence follows this interesting and rather important simplification: the intensity factor of energy alone determines whether a change will actually occur; while the capacity factor determines the quantity of change, or of work done. When the capacity factor is constant, or nearly so, more work can be got from a definite amount of energy with a high than with a low intensity factor, and a moment's reflection will show that in every transformation the intensity factor will be diminished. Energy becomes less available for doing work when the intensity factor is diminished.

What are the factors of chemical energy?—For, as A. Butterow said, in 1861, "it is necessary to distinguish the quantity of affinity from its intensity—i.e. the smaller or greater energy with which it tends to become active." If chemical energy can be resolved into two factors, the one factor must be analogous to the capacity, and the other to the intensity factor of thermal energy. J. W. Gibbs calls the intensity factor of chemical energy the chemical potential, and G. Helm calls it the chemical intensity, and it has been also called the driving force of a reaction. These terms are employed with the idea of avoiding the vagueness of the old term, chemical affinity, which is undoubtedly the correct designation for "chemical intensity." Now, the quantity of a substance which takes part in any chemical change is proportional to the "equivalent weights" of the substance, when the term "equivalent weight" refers to quantities of matter which have the same valency. Assuming that the chemical equivalent is the capacity factor of chemical energy, we may write—

Chemical energy = equivalent weight \times chemical affinity;

or

Chemical energy = equivalent weight \times chemical intensity.

Some follow M. Salet and write "valency" where "equivalent weight" has just been written; but that does not seem the right thing to do.

If two bodies at the same temperature be placed in contact, there will be no apparent conduction of heat from the one to the other; but when the temperature, or intensity factor, of the one body is higher than that of the other, heat will pass from the hot to the cold body, so that the cold body is warmed and the hot body is cooled. So with chemical energy. We assume that the molecules of every substance possess a specific amount of chemical energy, which has a definite intensity under certain specified conditions. One substance can only react with another when the intensity of the energy associated with the original mixture is greater than that of the final system. If the intensity of the energy associated with the original mixture be the same as that associated with the products of the reaction, no reaction will take place, for the system will be in stable equilibrium; if the intensity factors are not equal, the energy will not usually be at rest.

Water placed in a series of vessels in communication with one another will only come to rest when the surface of the water is at the same level in each vessel. "Difference of level" here means that the gravitational energy has a different intensity factor in each vessel. The availability, so to speak, of any given form of energy depends on the magnitude of the intensity factor. The real value of heat as a source of power depends on its temperature. The hotter a body in relation to its surroundings, the greater the available heat energy. This idea is involved in the so-called second law of thermodynamics. An electric current will flow whenever there is an inequality of the intensity factor—*i.e.* a difference of potential—at different parts of the circuit. If the intensity factors of any particular form of energy in a system are not equal, the system will be in a state of unstable equilibrium. Such a condition will not be permanent, and energy will flow, so to speak, from one part to another until the different intensity factors become equal.

Ostwald has drawn attention to the fact that if the chemical process be performed in a voltaic cell, the work derived from that process will be transformed into an equivalent amount of electrical energy. And since, by Faraday's law, the capacity factor—quantity of electricity—is proportional to the quantity of matter decomposed, the capacity factor of the electrical energy will be proportional to the capacity factor of the chemical energy. Hence the respective intensity factors of chemical and electrical energies will also be proportional. But electromotive force is proportional to the intensity factor of electrical energy, and therefore electromotive force is proportional to chemical affinity. Electromotive force and chemical affinity are manifestations of one form of energy. We see, then, with Faraday, that "the forces called electricity and chemical affinity are one and the same." Our problem is solved for conductors of electricity—electrolytes. Chemical action takes place when the potential of the reacting substances is greater than that of the reacting products. We can to-day express the "affinity" between a number of reacting substances roughly in terms of difference of potential. How this may be done for non-conductors of electricity has not yet been determined.

The temperature or intensity factor of heat energy required for the decomposition of many substances—say calcium or potassium chloride—is so great that commercial methods of decomposing these substances by thermal energy are not profitable. A great many compounds thus appear to be very stable when heated to high temperatures, these can often be decomposed by electrical energy at a comparatively low voltage (intensity factor). This illustrates how the commercial production of metals like aluminium, calcium, etc., were not particularly successful until electrical methods were adopted. The prediction of C. L. Berthollet, 1803, has been fulfilled. "The electric current," said he, "has furnished chemistry with an agent whose energy may be carried to a degree which, as yet, can scarcely be imagined, and which will furnish the means of producing in the formation and decomposition of chemical combinations, effects unforeseen, and superior to those which it is possible to obtain by the action of heat." It has been suggested, too, that if a source of energy with a particularly high intensity factor were available, it would most likely be possible to decompose many of the so-called elements into still simpler substances, but this, of course, is merely a speculation.

Questions.

1. What is the weight in grams of (a) iron, (b) aluminium which would theoretically be required to remove (*i.e.* precipitate) the copper completely from a solution containing one kilogram of crystalline copper sulphate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ($\text{Cu} = 63.5$, $\text{Zn} = 65$, $\text{Al} = 27$, $\text{O} = 16$, $\text{S} = 32$, $\text{H} = 1$) ?—*Sydney Univ.*
2. How much silver will be deposited from a solution of silver nitrate by a current of 2 ampères in 30 minutes ?
3. Explain the meanings of the terms "electropositive" and "electronegative" as applied to the elements.—*London Univ.*
4. What weight of silver and copper would be deposited by an electric current flowing through a solution of those metals in the same time that it liberates 30 c.c. (at n.p.t.) of a mixture of oxygen and hydrogen from acidulated water ($\text{Cu} = 63$, $\text{Ag} = 108$) ?—*London Univ.*
5. The solution pressures of the elements: manganese, zinc, tin, hydrogen, copper, and oxygen decrease in the order in which they are named. What might be expected to occur (and why) when (a) metallic manganese is placed in a solution of copper; (b) when metallic copper is placed in a solution of a tin salt; (c) when metallic manganese is treated with an acid; (d) when tin is treated with a weak acid; and (e) when copper is treated with an acid ?—*Massachusetts Inst. Technology.*
6. Show how the evolution of hydrogen when metallic iron is treated with hydrochloric acid is comparable with the deposition of lead upon a strip of zinc when the latter is immersed in a solution of a lead salt.—*Massachusetts Inst. Technology.*

CHAPTER XXI

COPPER, SILVER, AND GOLD

§ 1. Copper—Occurrence and Properties.

History.—Copper appears to have been known from prehistoric times—the neolithic age—long before the histories of the ancient peoples were engraved on stone pillars or inscribed on papyri. The “copper age” followed the “stone age.” At the time of Ramses II (about 1300 B.C.), copper was so costly that it was stored among the treasures of Egyptian temples. Copper appears to have been used for making utensils and instruments for war before iron. This is probably due to the fact that copper occurs native in a form requiring no metallurgical treatment. The ancients used the terms $\chi\alpha\lambda\kappa\acute{o}\varsigma$ (chalcos) and $\alpha\epsilon\varsigma$ for copper, brass, and bronze. In fact, the terms for copper, brass, and bronze are much confused by the old writers, showing that they did not understand the difference—e.g. Pliny. Pliny, by the way, wrote a celebrated *Natural History* at the beginning of this era. In this book he laboriously garnered what he could of facts and fable then known concerning nature; hence the book is often quoted. Copper was afterwards called $\alpha\epsilon\varsigma$ *cyprium* (i.e. Cyprian brass), since the Romans first obtained it from the Isle of Cyprus; the term $\alpha\epsilon\varsigma$ *cyprium* was soon abbreviated to *cuprum*. Hence the modern symbol “Cu.” The seven metals—gold, silver, mercury, copper, tin, iron, and lead—known to the earlier chemists were designated by the names and symbols of the seven greater heavenly bodies—Sun, Moon, Mercury, Venus, Jupiter, Mars, and Saturn. Thus the looking-glass of Venus ♀ symbolized copper. In some cases it is possible to see a reason why a particular metal was assigned to a particular heavenly body, but in other cases the connection is more remote.

Occurrence.—Metallic copper is found in many localities; e.g. considerable masses have been found in Michigan, on the shores of Lake Superior; and small quantities in many other places—Cornwall, Siberia, Ural, Australia, Chili, etc. Compounds of copper are distributed in nature as oxide in *cuprite*, or ruby ore, Cu_2O ; as sulphide in *chalcocite*, or copper glance, Cu_2S ; copper pyrites or *chalcopyrite*, CuFeS_2 , or $\text{Cu}_2\text{S} \cdot \text{Fe}_2\text{S}_3$. The real composition of many copper sulphides, as they occur in nature, is exceedingly complex. The same remark is more or less true for the composition of most natural minerals; at least chemical formulæ which follow the analyses closely are very complex. The formulæ for minerals are commonly represented as if pure minerals occurred in nature. Ideally pure minerals very seldom occur in nature, and accordingly, the

formulae represent ideal or imaginary minerals to which real minerals approximate more or less closely. Copper also occurs in many places as basic carbonate, *malachite*, $\text{CuCO}_3 \cdot \text{Cu(OH)}_2$; and *azurite*, $2\text{CuCO}_3 \cdot \text{Cu(OH)}_2$. Copper silicates, arsenates, phosphates, etc., are also known to occur native. Copper has also been found in the feathers of some birds.

Properties.—Copper has a characteristic reddish-brown colour by reflected light, but in transmitted light, thin layers are green. The metal can be obtained in octahedral crystals (cubic system). When near its melting point, copper is brittle enough to be powdered. If a piece of copper be heated and cooled slowly, it is brittle; and if cooled rapidly, it is soft, malleable, and ductile. Copper is one of the best conductors of heat and electricity. These properties are modified if the metal be contaminated with minute traces of impurity. Copper melts at $1083\cdot0^\circ$ in an atmosphere of carbon monoxide; if heated in air, the melting point is reduced some 20° . This is due to the formation of cuprous oxide which dissolves in the molten metal. Copper begins to volatilize at about 900° , and it can be distilled in the electric furnace. It furnishes a green vapour, and it colours Bunsen's flame green. Dry air has no action on copper, but in the presence of atmospheric moisture and carbon dioxide, the metal becomes covered with a green basic carbonate called *verdigris*—French *vert de gris*. The atmosphere of towns containing sulphur oxides may also form basic sulphates with the copper. Hydrochloric and sulphuric acids have little or no action on the metal in the cold; hot sulphuric acid dissolves copper rapidly, and hot hydrochloric acid attacks the metal slowly. Nitric acid—hot or cold, dilute or concentrated—dissolves the metal rapidly, forming copper nitrate. Ammonia acts on copper in the presence of air, forming a deep blue solution.

Atomic weight.—The combining weight of copper has been determined by the analysis or synthesis of the oxide, sulphate, chloride, etc. The results, referred to oxygen = 16, vary between 63·35 and 63·68; and the best representative value is taken to be 63·57. This number also represents the atomic weight, as estimated by the isomorphism of the cuprous, silver, and gold salts; by the vapour density of the volatile copper compounds; and by Dulong and Petit's method of approximation—specific heat of copper 0·0956.

Uses.—Next to iron, copper is the most useful metal. Enormous quantities are used in the electrical industries. It is also made into household utensils, boilers, etc. Copper nails, rivets, and sheeting are used for sheathing ships because copper is but slowly corroded in moist air and in sea-water. Copper is one of the chief ingredients in numerous industrial alloys. A high degree of purity is not always a desideratum, since the pure metals occasionally possess undesirable qualities. Copper is one of the few metals which in the pure state has an important field of usefulness, since it is then a far better conductor of electricity than the less pure metal. Gold and silver, used in jewellery, etc., are usually alloyed with other metals—e.g. copper—in order to give them hardness and durability without detracting from their appearance. New alloys of the various metals are constantly being investigated. *British copper coins* contain 95 per cent. of copper, 4 per cent. of tin, and 1 per cent. of zinc; part of recent issues had copper 95·5, tin 3·0, and zinc 1·5 per cent. Gold and silver coins of different nations usually contain 8–10 per cent. of copper.

Nickel coins in Germany and the United States contain about 25 per cent. nickel, and 75 per cent. of copper. Copper is largely used in the manufacture of alloys. With zinc it forms *brass* (zinc 1, copper 2-5)—*common brass* has zinc 1, copper 2, *Dutch metal* (zinc 1, copper 4), *bell metal* (copper 3, zinc 1) : with tin, *speculum metal* (tin 1, copper 2) used for optical instruments; *gun metal* (tin 1, copper 9) was once used for making cannon. *Bronze* is an alloy of copper (70-90 per cent.), zinc (1-25 per cent.), and tin (1-18 per cent.); it is used for making statues, coins, ornaments, etc. *Phosphor bronze* contains tin and a small percentage of phosphorus; *manganese bronze* employed for propeller blades has about 0.3 per cent. of manganese. The manganese alloys are tougher than ordinary bronze, and they are largely used for steamship propellers because they resist corrosion by sea water fairly well, and certain parts of machinery. *Aluminium bronze* contains 90 per cent. of copper; it is a hard yellowish-brown alloy, light, strong, and elastic. It is used for making the hulls of yachts, etc. *German silver* contains copper (56-60 per cent.), zinc (20-25 per cent.), and nickel (20-25 per cent.). It is used in making resistance coils (owing to its low electric conductivity), and for imitating silver.

§ 2. Copper—Extraction.

The methods employed for winning copper from its ores depend upon the kind of ore used, and upon local conditions. Similar ores are seldom treated in the same manner in different localities. If the ore contains no sulphur—as is the case with the carbonates and oxides—the metal can be obtained by simply reducing the ore mixed with suitable fluxes and coke in a furnace heated by a blast. Assuming that the ore to be smelted is a typical chalcopyrite, the operations involved in extracting the copper are somewhat complex because (1) a large number of impurities are present in the ore; (2) copper of a high degree of purity is needed; and (3) it will probably be profitable to recover gold and silver from the crude metal. Nearly all the copper produced to-day is made in a modified form of Bessemer's converter.

1. *Roasting the ore.*—The pyritic ore is first crushed, and it may or may not be necessary to concentrate the pyrites by washing away the rocky impurities. The concentrated ore, in one works, contained about 14 per cent. of copper, and 29 per cent. of iron. Part of the concentrated ore is then roasted, say, in a reverberatory furnace in order to convert the sulphides into oxides. The copper sulphide oxidizes less rapidly than iron sulphide; such is the difference, that if a mixture of the two sulphides be acted upon by air, it is usually stated that "the copper sulphide is not attacked until all the iron sulphide is oxidized." The reaction in extreme cases is represented: $2\text{CuFeS}_2 + 6\text{O}_2 = \text{Cu}_2\text{O} + \text{Fe}_2\text{O}_3 + 4\text{SO}_2$. It may or may not be convenient to use the sulphur dioxide for the manufacture of sulphuric acid.

2. *Fusion for matte.*—A mixture of roasted and unroasted ("green") ore and coke is charged into a reverberatory furnace or into a blast furnace lined with firebricks and heated by a blast. The ore probably contains silica, if not, some must be added when the furnace is charged. The air-blast burns the carbon to carbon monoxide: $2\text{C} + \text{O}_2 = 2\text{CO}$. Part

of the cuprous oxide formed in the preceding operation is reduced to copper by the joint effect of the carbon and carbon monoxide: $\text{Cu}_2\text{O} + \text{CO} = 2\text{Cu} + \text{CO}_2$. The copper oxide is much more readily reduced than iron oxide. The copper unites with the sulphides in the unroasted ore: $3\text{Cu} + \text{Fe}_2\text{S}_3 = 3\text{CuS} + 2\text{Fe}$; and some of the unreduced cuprous oxide forms cuprous sulphide: $3\text{Cu}_2\text{O} + \text{Fe}_2\text{S}_3 = 3\text{Cu}_2\text{S} + \text{Fe}_2\text{O}_3$; and some cuprous sulphide and oxide react to form metallic copper and sulphur dioxide: $2\text{Cu}_2\text{O} + \text{Cu}_2\text{S} = \text{SO}_2 + 6\text{Cu}$. Any cupric sulphide present is reduced to cuprous sulphide. The net result of the smelting for matte is (1) the transformation of the copper into cuprous sulphide, (2) part of the iron unites with the silica to form a fusible slag, and (3) part is reduced to ferrous sulphide which remains admixed with the cuprous sulphide to form matte. The lighter slag rises to the surface and flows over into a pit. The matte, which collects at the bottom of the trough, is drawn off from time to time. Matte is a more or less impure mixture of cuprous and ferrous sulphides containing 20–75 per cent. of copper. The gold and silver, and part of the arsenic and antimony in the ore remain with the matte. The furnace is then tapped, and the matte and slag are run into a trough.

3. The conversion of the matte into blister copper.—The molten matte is run into a tilting “converter” furnace lined with a mixture of quartz

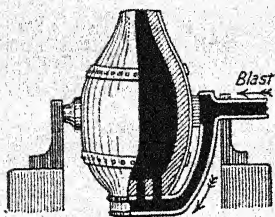


FIG. 158.—“Converter” Furnace (Diagrammatic).

and clay, and arranged with openings in the bottom so that air can be blown through the molten matte, Fig. 158. The sulphur, iron, and many other metals are oxidized. The volatile oxides are driven off, the iron oxide unites with the siliceous lining of the converter to form a slag. The iron and sulphur are oxidized first. As soon as appreciable quantities of copper commence to oxidize, the operator can tell from the appearance of the flame issuing from the converter that it is time to stop the blast. The furnace is then tilted, and

the copper is run off. As the copper cools, the sulphur dioxide dissolved by the metal is expelled, this gives the metal a blistered appearance, hence the product is called blister copper.

4. The conversion of blister into crude copper.—The blistered copper is melted on the hearth of a reverberatory furnace—e.g. Fig. 192—and any oxide dissolved by the copper is reduced by forcing a log of green wood into the molten mass. The hydrocarbons of the wood, bubbling through the molten metal, reduce copper oxide to metallic copper. The “poled” copper is cast into plates—about $\frac{3}{4}$ -inch thick, 3 feet wide, and 3 feet long—if it is to be further purified.

5. Refining the crude copper by electrolysis.—Crude copper is refined by an electrolytic process. Plates of crude copper are suspended, as anodes, in a bath of copper sulphate acidified with sulphuric acid. Sheets of pure copper, as cathodes, are suspended alternately with the anodes in the same bath. When the current passes, copper dissolves from the anodes, and pure copper is deposited at the cathodes. The impurities in the crude copper either pass into solution, or are deposited as a mud or “slime”

about the anode. The anode is enclosed in filter cloth bags to facilitate the collection of the "anode mud." Considerable amounts of silver and gold are obtained from the slimes or "anode mud." It is possible to get a very high degree of purity—e.g. 99.8 per cent. copper—by the electrolytic process. This quality of copper is needed for electrical purposes, since a trace of impurity may considerably reduce the electrical conductivity of the metal. Copper matte, blister copper, and even copper ores have been treated, more or less successfully, by electrical processes, but the product is not very pure.

§ 3. Silver—Occurrence and Properties.

History.—Silver has been known from ancient times. There are some allusions to silver in the Old Testament, and it was probably used as money as early as gold. The Phœnicians are supposed to have obtained their silver from Armenia and Spain. Silver appears to have been purified by a process of cupellation, but there is little evidence to show that the ancients knew how to separate silver from gold. The old terms for silver refer to its bright white colour—the Hebrew equivalent is derived from the verb "to be white," and the Greek term from ἀργός (argos) to be shining. The early chemists termed silver "Luna," or "Diana," and represented it by the symbol ☾ for the crescent moon—probably because of the pale silvery colour which the poets say is characteristic of moonlight.

Occurrence.—Native silver is occasionally found in large masses or crystallized in cubes or octahedra. It is also found associated with metallic copper, gold, etc. The principal ores of silver contain silver glance or *argentite*, Ag_2S , admixed with several other sulphides—antimony, arsenic, and copper. The chief silver ores are found in Mexico, Peru, Chili, Bolivia, Idaho, Arizona, Nevada, Colorado, Norway, Australia, etc. Much silver also occurs associated with lead in galena, and a great deal of the silver in commerce is extracted from argentiferous lead. Silver chloride, AgCl , occurs as *kerargyrite*, or horn silver.

Properties of silver.—Silver is a white lustrous metal which appears yellow if the light be reflected from its surface many times before it reaches the eye. Very thin layers of silver have a bluish tint. Powdered silver is grey and earthy in appearance. Silver is highly malleable and ductile. Sheets 0.00001 inch thick have been made. Silver melts at 958° in an atmosphere of carbon monoxide, and vaporization commences about 850° and proceeds rapidly between 1200° and 1500° . It can be distilled in the oxy-hydrogen blowpipe, or in the electric furnace. The vapour appears of a greenish colour. Molten silver absorbs oxygen, and rejects most of it on cooling. The evolution of oxygen from cooling silver is often violent enough to spurt some of the metal away from the main mass—hence the term *the spitting of silver*. Silver conducts heat and electricity better than copper. Exposure to the air has no action on silver, but if the air be contaminated with hydrogen sulphide, the silver is blackened owing to the formation of a thin film of silver sulphide on the surface. Nitric acid—hot or cold, dilute or concentrated—readily dissolves the metal, forming silver nitrate. Hot concentrated sulphuric acid gives silver sulphate, but

the metal is not perceptibly attacked by the dilute acid. Hydrochloric acid acts very slowly, if at all, at ordinary temperatures, but at a red heat, hydrogen chloride forms silver chloride. The vapour density of silver vapour at 2000° shows that at this temperature the molecules are monatomic.

Atomic weight.—Analyses of various halogen compounds of silver give numbers ranging from 107·67 to 108·09 for the combining weight of silver (oxygen = 16); the best representative value is supposed to be 107·88 which also represents the atomic weight. This number agrees with the isomorphism of the silver, gold, copper, and sodium salts; and also with Dulong and Petit's method of approximation—specific heat of silver, 0·05625.

Uses.—British silver coinage has very nearly 92·5 per cent. of silver and 7·5 per cent. of copper. This is the standard of *sterling silver* for coinage and for silver plate regulated by law down to 1919; since 1920, the silver coinage has 50 per cent. of silver. In 1920–22 the silver coinage had 41 per cent. of copper, and 9 per cent. of nickel, and later issues have had 50 per cent. of copper. The American dollar has about 90 per cent. of silver. Silver ornaments made from standard silver can be heated in air to oxidize the copper near the surface of the metal, the resulting copper oxide is removed by digestion with sulphuric acid, leaving a superficial layer of pure silver. The effect so produced is called “frosted silver.” The so-called “oxidized silver” is made by dipping silver ornaments in a solution of an alkaline sulphide. A thin film of sulphide is thus formed on the surface.

§ 4. Silver—Extraction.

Although the methods employed for the extraction of silver are very varied, yet they can be arranged into four groups:

1. **Lixiviation processes.**—In Ziervogel's process, the ores are roasted under carefully regulated conditions so that the iron and part of the copper sulphides are converted into oxides, while the silver and part of the copper are converted into silver and copper sulphates, Ag_2SO_4 , CuSO_4 . The soluble sulphates are extracted with water, and the silver precipitated from the solution by the addition of scrap copper. The copper is afterwards precipitated by the addition of iron. This process was once much used for argentiferous copper mattes. In Percy and Patera's process the ore is roasted with salt; silver chloride is formed. This is extracted with sodium thiosulphate; or with strong brine, as in Augustin's process. In Percy and Patera's process the silver is precipitated from the solution as silver sulphide by the addition of sodium or calcium sulphide and reduced to the metal by calcination in a roasting furnace; in Augustin's process, the silver is precipitated by scrap copper. The modern cyanide process has practically ousted the lixiviation processes just indicated. In the cyanide process, the crushed ore is treated with a solution of potassium or sodium cyanide. The cyanide reacts with the silver sulphide: $\text{Ag}_2\text{S} + 4\text{NaCy} \rightleftharpoons 2\text{NaAgCy}_2 + \text{Na}_2\text{S}$. The accumulation of sodium sulphide in the solution stops the reaction. When the solution is exposed to the air, however, the sodium sulphide is oxidized to sodium thiosulphate and sulphur. The reaction then progresses as indicated in the equation from left to right. Thus, the free access of air to the cyanide

solution is an important factor in promoting the dissolution of silver. The silver is recovered from the solution by zinc or aluminium precipitation.

2. Amalgamation processes.—In Mexico, where fuel is scarce, the so-called patio process—invented about 1557 by B. de Medina—has been in use over 300 years, but now it is almost displaced by the cyanide process. Hence the patio process promises very soon to be of little more than historic interest. In this process, the powdered ore mixed with water is thoroughly incorporated with a little sodium chloride. In about a day, mercury is added, along with some roasted pyrites containing a mixture of copper and iron sulphates and oxides. The whole is most intimately mixed, and more mercury added from time to time. The mixing is done on large stone-paved areas—patios—by being trodden by mules. It is probable that copper chlorides are first formed: $2\text{NaCl} + \text{CuSO}_4 = \text{CuCl}_2 + \text{Na}_2\text{SO}_4$, and these decompose the silver sulphide: $\text{CuCl}_2 + \text{Ag}_2\text{S} = 2\text{AgCl} + \text{CuS}$; $\text{CuCl}_2 + \text{CuS} = 2\text{CuCl} + \text{S}$; and $2\text{CuCl} + \text{Ag}_2\text{S} = 2\text{AgCl} + \text{Cu}_2\text{S}$. The silver chloride dissolves in the brine, and it is then immediately reduced and dissolved by the mercury: $\text{AgCl} + \text{Hg} = \text{HgCl} + \text{Ag}$. The mercury is more active if about 1 per cent. of sodium be dissolved in it. The resulting silver amalgam is separated from the mud by washing and settling; and the excess of mercury is squeezed through canvas bags. The solid amalgam is then pressed into cakes, and heated in a retort. Mercury distils over, and silver remains behind. A certain amount of mercurous chloride is lost during the operations.

3. Smelting process is largely used in America. The object is to concentrate the silver in lead. The ore is mixed with coke and limestone, and heated in a small blast furnace. A fusible silicate of iron and lime is formed as a slag, and the reduced metal accumulates in the bottom of the furnace. Slag and metal are run off from time to time. The “pigs” of lead are passed on for desilverization.

4. Electrolytic processes.—The electrolytic process applied in the refining of copper extraction is used for the separation of silver from argentiferous copper; and it would be difficult to find a more remarkable application of the electric current by transferring the metal to be purified from anode to cathode. In B. Moebius' process (1884), the electrolyte is a mixture of silver and copper nitrates acidified with nitric acid. A slab of impure silver is used as anode, and a plate of pure silver as cathode. Silver and copper are dissolved at the anodes, and silver is precipitated on the cathodes. The gold, if present, remains undissolved as a slime below the anode. The anodes are enclosed in filter bags to facilitate the collection of the “anode mud.” The composition and concentration of the solution, as well as the current density are carefully regulated on account of the danger of depositing copper with the silver.

§ 5. Desilverization of Lead.

Lead can be desilverized by means of Pattinson's or Parkes' process. The latter process is generally used, the former but seldom.

1. Pattinson's process.—Molten lead and silver mix together in all proportions. The melting points of all possible alloys of the two metals are indicated in Fig. 159. There is a eutectic at 303° , when the molten

mixture contains $2\frac{1}{2}$ per cent. of silver. If a molten mixture of lead with, say, 50 per cent. of silver be allowed to cool, when the temperature reaches 648° , some silver (contaminated with a little lead) will separate from the solution. The remaining fluid is richer in lead than it was before, and consequently remains fluid, but as its temperature falls, more silver will separate from the solution and the mother liquid will continually get poorer and poorer in silver until but $2\frac{1}{2}$ per cent. remains. The whole will then freeze *en masse* at 303° . If the molten mass contains less than $2\frac{1}{2}$ per cent. of silver, it will begin to freeze somewhere between 327° and 303° , and lead (contaminated with a little silver) will separate from the

solution, and continue separating until the mother liquid has $2\frac{1}{2}$ per cent. of silver, when all will freeze *en bloc*. The results are here quite analogous with the freezing of solutions of salt and water indicated in an earlier chapter. In Pattinson's process for separating silver from lead, the molten mixture containing less than $2\frac{1}{2}$ per cent. of silver is allowed to cool slowly, and the crystals which separate from the cooling solution, containing more lead than the original solution, are skimmed off with perforated ladles

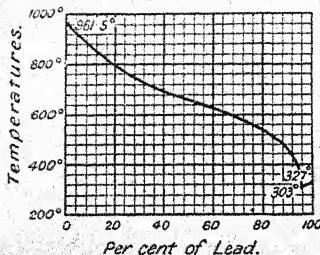


FIG. 159.—Freezing Curve of Silver: Lead Alloys.

into a neighbouring pot until one-third to one-eighth of the original solution remains. The enriched silver-lead alloy remains in the pot. The operations are repeated on both fractions until the enriched lead contains about one per cent. of silver, and the desilverized lead contains but 0.001–0.002 per cent. of silver. The theoretical enriched lead should have $2\frac{1}{2}$ per cent. of silver. The enriched lead is then cupelled for silver—see “Gold refining by cupellation.”

II. Parkes' process.—Zinc readily forms a number of chemical compounds with silver—e.g. Ag_3Zn_2 , AgZn , Ag_2Zn_3 , Ag_5Zn_3 , AgZn_{12} . They all show maximum points on the freezing-point curve of alloys of zinc and silver. These compounds all freeze at a higher temperature than molten lead, and they are specifically lighter than molten lead. Hence if molten zinc be stirred in a bath of molten argentiferous lead, and the molten mixture be allowed to cool, a crust is formed on the surface containing a zinc-silver compound. Gold in the argentiferous lead also passes with the silver to the zinc. The zinc-silver crystals are removed from the surface as they are formed until the lead itself begins to solidify. The skimmings can be remelted, and the process repeated, until finally the remaining lead contains but 0.0005 per cent. of silver. The desilverization can thus be carried much further by this process than is practicable with Pattinson's. In practice, the zinc added to the lead is kept between 0.8 and 1.5 per cent., and the zinc-silver crystals which separate are mainly Ag_5Zn_3 . The mixture of zinc, lead, etc., is heated on an inclined surface hot enough to melt the lead, but not the zinc alloy. The lead flows away, and the enriched scum remains. The zinc can be separated from the silver, gold, and lead, by distillation. The residue in the retort containing the silver and gold is cupelled.

§ 6. Gold—Occurrence and Properties.

The extensive use of this metal wherever gorgeous ornament is required, as well as in the simplest designs of art—from the glittering crown of an emperor to the neat little wedding-ring of the village maid—has imparted to gold an interest which other metals have failed to excite.—H. SOWERBY.

History.—"Gold," said T. Bergman, "occurs in the bowels of the earth native, possessing a complete metallic form, although in general the small particles of it are so interspersed in various matrices that they are entirely invisible." Gold must have been one of the earliest of metals to attract the attention of primitive man, since it occurs free as virgin gold in nature, and is found in the rocks and gravels of many rivers. Flint daggers with gilt handles have been reported from excavations in Egypt, and gold is mentioned in the earliest writings of civilized man. Representations of quartz-crushing and gold-refining processes are reported to have been found in Egyptian tombs dated 2500 B.C.; similar remarks apply to a map showing the gold-mining regions 1350-1330 B.C. The gold mines of Nubia were worked extensively by the early Egyptians. Pliny described the amalgamation process for the extraction of gold. Cupellation processes for the purification of gold were described in the second century, and the same process is probably referred to by Jeremiah in the Old Testament, 600 B.C. The earliest words for gold in Hebrew and Egypt refer to its shining appearance. The early chemists called gold *sol* (the sun) or Apollo, and represented it by the symbol of perfection ☉ or ✱ or by the sun or Apollo with a crown of rays on his head, Fig. 1; not altogether on account of its appearance, but because they considered it to be the most perfect of the noble metals—*metallum rex*, the very king or Apollo of metals.

Occurrence.—Gold is generally found in a metallic condition in quartz veins as reef gold; and in alluvial gravels as alluvial gold—the latter represents the *débris* from the weathering of auriferous rocks which has been washed into river beds, etc. Large nuggets are occasionally found—one from California weighed over 190 lbs.; and one from Victoria, 183 lbs. Native gold is never found pure, but specimens 99 per cent. purity are sometimes found; and one from Cripple Creek (Colorado) was reported to be 99.9 per cent. purity.

Metallic gold is very widely distributed in nature in quantities too small to be profitably extracted. Sea-water, for instance, is said to contain about $3\frac{1}{2}$ grains per ton. Granite, on the average, has about 0.37 part per million; sandstones, 0.03 part per million; limestones, 0.007 part per million. Gold also occurs in small quantities in clays, iron pyrites, and in almost all silver, copper, bismuth, lead, zinc, tellurium, and antimony ores. Gravels which need not be crushed can sometimes be profitably treated for gold—alluvial gold—if but 2 to 3 grains per ton be present, that is, one part of gold per 5 million parts of worthless material. The mean of the returns for the Rand is something less than half an ounce of gold per ton of material treated.

Properties.—Gold! Yellow glittering precious gold is yellow only when in masses, for it appears red if the light be reflected many times from the surface of the metal before it reaches the eye. Gold-leaf is green or blue in transmitted light; this is best seen by placing a sheet of leaf-gold

between two sheets of glass and examining the combination through a magnifying-glass in transmitted light. If precipitated in a fine state of subdivision, the tint of gold varies from red to dark brown. Very thin films of gold are crimson or purple in transmitted light. Gold crystallizes in the cubic system. It is one of the most malleable and ductile of metals, sheets 0·0000004 inch thick have been made. It is not quite so good a conductor of heat and electricity as silver and copper. Gold melts at 1063·0°; and the molten metal appears green. It begins to volatilize at temperatures just below its melting point, say, at 970°. Krafft and Bergfeld say that the metal boils at 2530°. Gold occludes oxygen, hydrogen, and carbon monoxide under suitable conditions. Gold is not acted upon by air or oxygen at any temperature, hence the alchemists called gold a *noble metal* in contrast with *base metals*—like copper, lead, tin, etc.—which are oxidized and lose their metallic character when heated in air. Silver and platinum are noble metals for the same reason as gold. The inert gases argon and its congeners have been called *noble gases* because they are chemically inactive. The fact that Moses was able to reduce Aaron's "golden" calf to ashes by calcination referred to in the book of Exodus is taken as evidence that the gold of Aaron and his religious brethren was a kind of brass, possibly alloyed with a small proportion of gold.

Solvents for gold.—Gold is scarcely affected by nitric, sulphuric, and hydrochloric acids, but it is dissolved by aqua regia—i.e. a mixture of nitric and hydrochloric acid—or by a mixture of hydrochloric acid with an oxidizing agent which liberates chlorine; by water containing the halogens chlorine, bromine, or iodine in solution; and by solutions which can generate the halogens. The action of aqua regia on gold is sometimes represented in symbols: $\text{Au} + \text{HNO}_3 + 3\text{HCl} = \text{AuCl}_3 + \text{NO} + 2\text{H}_2\text{O}$. The alchemists were wont to represent the dissolution of this metal by some fierce animal devouring the sun or Apollo—Fig. 1. Gold is attacked by boiling ferric chloride solutions, hot selenic acid, telluric acid with sulphuric or phosphoric acid, alkaline sulphides and thiosulphates, perchlorates, perbromides, and periodides of the metals, iodic and periodic acids with hot sulphuric acids, and by reacting substances which give large quantities of oxygen—manganese dioxide or potassium permanganate or nitric acid with sulphuric acid—and aqueous solutions of potassium cyanide when exposed to the air. Gold is not appreciably attacked by solutions of the alkalis. The freezing point of a solution of gold in mercury corresponds with a monatomic molecule Au.

Atomic weight.—Analyses of gold chloride, gold bromide, and some of the aurates furnish values for the combining weight between 197·05 and 198·25; the best representative value appears to be 197·2. This number probably represents the atomic weight judging from the isomorphism of the silver, cuprous, aurous, and sodium salts; and from Dulong and Petit's method of approximation—specific heat of gold, 0·0316.

Uses.—British gold coinage contains 91·66 per cent. of gold and 8·33 per cent. of copper. This is called *sterling or standard gold*. The gold coins usually have 7–8 per cent. copper, and the remainder is silver. Gold alone is too soft to withstand wear and tear. The gold coinage of Sydney mint has the same amount of gold, but silver is used in place of copper, so that the Sydney sovereign is greenish-yellow. The amount of gold

in alloys is usually expressed in terms of the *carat*. Here, 1000 parts are divided into 24 equal parts. Pure gold is 24-carat gold. The sovereign is a 22-carat alloy because it contains 22 parts of gold per 24 parts of other metals. The standard gold alloys recognized by law are 22-, 18-, 15-, 12-, and 9-carat, or parts of gold per 24 parts of alloy. Ware made by these standard alloys can be "hall-marked." American gold coinage has 90 per cent. of gold and 10 per cent. of copper.

§ 7. Gold—Extraction and Refining.

The old historian Strabo tells us that the ancients washed gold down from the hills by torrents of water, and received it in sieves and sheepskins. This is supposed to be the source of the mythical expedition of the Argonauts under Jason in quest of the golden fleece. The clumsy methods of extraction by individual workers were more or less hit or miss, since they were successful in dealing with comparatively rich patches. Gold mining was therefore more or less of a gamble. To-day it is a carefully organized industry with approximately the same output from day to day.

Washing processes.—The amount of gold in a ton of ore is small, and in consequence, relatively large quantities of ore have to be treated. Gold is separated from alluvial gravel by mechanical washing. The specific gravity of gold is so much greater than that of the associated materials, that, when the mixture of sand and gravel is agitated with water in large pans or "cradles," and the rocky matters floated off, the fine particles of gold remain on the bottoms of the cradles as "gold dust"—panning or pan washing. This primitive method of washing has been replaced by placer mining, in which the sand containing the gold is agitated in sluices, that is, in long flumes or troughs with transverse cleets along the bottom, and through which powerful streams of water flow. The water sweeps away the sand, and the heavier gold collects on the bottom of the sluices. In hydraulic mining, water under high pressure is directed against the "earth" containing the gold. The "earth" and gold are washed into the sluices as in placer mining. Dredging is also used for, say, river-bottoms in order to get auriferous sands into sluices. In vein mining, the gold-bearing quartz is mined by blasting, etc.; the quartz is crushed to fine powder in stamper mills, and the gold extracted by one of the following processes:

Amalgamation process.—The powder from the stamper mills is ground in the so-called tube-mills—i.e. revolving cylinders lined with quartz blocks, and containing flint or quartz pebbles. The ground ore is floated as a "slime" in a large trough through which water is continually flowing. The slime then runs over copper plates amalgamated with mercury and kept in constant motion—*pulsator table*. The particles of gold amalgamate with the mercury. After some time the plates are scraped. The mercury is squeezed from the scrapings, and separated from the amalgam by distillation. Gold remains behind in the retorts. This is then fused and cast into ingots to be subsequently refined. Some gold still remains in the *slimes, tailings, or battery pulp* which have passed over the copper plates, and can usually be profitably treated by the cyanide process, and more gold extracted.

Chlorination process.—Gold can be extracted from pyrites by the chlorination process. The ore is first roasted, then wetted, and exposed in revolving barrels to the action of chlorine gas. The gold forms auric chloride, AuCl_3 , which is extracted by water, and precipitated from the solution by ferrous sulphate, or by hydrogen sulphide—followed by roasting of the precipitated sulphide.

Cyanide process.—The powdered ore is leached with a dilute aqueous solution of potassium cyanide (0.25 to 1 per cent.) while freely exposed to the atmospheric air. Under these conditions, gold is dissolved as a double cyanide: $4\text{Au} + 8\text{KCy} + 2\text{H}_2\text{O} + \text{O}_2 = 4\text{KOH} + 4\text{KAuCy}_2$. Gold is precipitated from the solution either by the addition of zinc shavings: $2\text{KAuCy}_2 + \text{Zn} = \text{K}_2\text{ZnCy}_4 + 2\text{Au}$; or by electrolysis. The gold is collected as a compact mass by fusion with sodium carbonate and charcoal.

Gold Refining.

Gold is refined by one of the following processes:

1. **Electrolysis.**—The anode is the alloy to be purified; the electrolyte is a solution of gold chloride in hydrochloric acid, and the cathode is pure gold. On electrolysis, fairly pure gold is deposited on the cathode, silver forms silver chloride which remains as a deposit about the anode.

2. **Cupellation processes.**—Gold is alloyed with an easily oxidizable metal, say, lead. The alloy of lead and gold is heated in a stream of air in a furnace with a shallow hearth made of bone ash. The lead is oxidized to litharge, PbO , which is then partly blown from the surface of the molten metal, and partly absorbed by the bone ash. When the gold appears as a bright disc, the operation is stopped and the gold removed. If silver be present, silver and gold remain alloyed after the operation, and they must be separated by some other process—say “parting with sulphuric acid.”

3. **Parting with sulphuric acid.**—In the old process of parting silver and gold, known as quartation, an alloy of gold and silver, containing less than 25 per cent. of gold, was treated with nitric acid. The silver dissolved as silver nitrate, and the gold remained behind undissolved as a brown powder. Parting with sulphuric acid is cheaper. Here the alloy is boiled with concentrated sulphuric acid, then with nitric acid. The gold is not attacked, and it remains behind as a brown porous mass. This is washed, dried, and fused into a compact mass with sodium carbonate and charcoal.

§ 8. Colloidal Gold and Silver.

If a solution containing 0.01 to 0.001 per cent. of gold chloride be made slightly alkaline by the addition of magnesia, and then a few drops of a reducing agent—formaldehyde, oil of turpentine, aqueous solution of acetylene, solution of phosphorus in carbon disulphide, carbon monoxide gas, sodium hyposulphite, $\text{Na}_2\text{S}_2\text{O}_4$, etc.—be added, the solution will probably acquire a ruby-red colour. Metallic gold is present in the form of minute particles which do not settle under the influence of gravity, and consequently remain suspended an indefinite time, and the solution can be filtered through paper unchanged. The solution is sometimes

called Faraday's gold, because it was studied by M. Faraday in 1857. The gold is in colloidal solution because the clear solution appears opalescent by Tyndall's optical test, and the particles can be perceived by the ultra-microscope. Similarly, by the action of certain reducing agents on soluble silver salts, Carey Lea (1887) obtained coloured solutions containing colloidal silver from which finely divided silver was obtained coloured golden-yellow, ruby, blue, etc.

The metallic gold can be removed from its colloidal solution by shaking the solution with precipitated aluminium hydroxide, stannic hydroxide, barium sulphate. The decolorization here resembles the decoloration of coloured solutions by shaking them with recently ignited charcoal. The gold is absorbed by the precipitating agent. Animal charcoal, if shaken with the solution, adsorbs the gold. The addition of electrolytes—acids, neutral salts, and alkalies—changes the red colour to blue, then violet, and then black. This is due to the coagulation of the particles of gold into clots. The gold then settles to the bottom of the fluid.

When an insoluble precipitate is formed in the absence of electrolytes by a reaction between two chemical compounds, it is almost always in the colloidal condition. Thus if aqueous hydrogen sulphide be added to a solution of arsenious acid, a turbid yellow solution of colloidal arsenic sulphide is formed: $2\text{H}_3\text{AsO}_3 + 3\text{H}_2\text{S} \rightleftharpoons \text{As}_2\text{S}_3 + 6\text{H}_2\text{O}$. If the precipitate be made by adding an aqueous hydrogen sulphide to a solution of arsenious chloride, a coagulated precipitate of arsenious sulphide is formed. In the latter case, hydrochloric acid is produced by the reaction: $2\text{AsCl}_3 + 3\text{H}_2\text{S} \rightleftharpoons \text{As}_2\text{S}_3 + 6\text{HCl}$. If some hydrochloric acid be added to colloidal arsenic sulphide formed as just indicated, the suspended colloid is at once coagulated and precipitated. These facts illustrate a principle of great importance in quantitative analysis where successful work depends upon the formation of an insoluble precipitate which can be easily washed free from absorbed mother liquid. When a colloid is precipitated by an electrolyte, as when aluminium and ferric hydroxides are precipitated by ammonia in the presence of ammonium chloride, the salt, ammonium chloride, coagulates the colloidal hydroxides into the gel condition. During the washing of the gel precipitate, the gel passes into the sol condition, because the coagulating salt is removed by washing. Hence a solution of ammonium nitrate is used for washing aluminium and ferric hydroxide precipitates in order to keep the colloid in the coagulated or gel condition. The ammonium nitrate is driven off during the ignition of the hydroxides. Gel colloids are said to be reversible colloids when they can be converted into the sol condition by restoring the original conditions, e.g. by washing out the coagulating electrolyte from precipitated aluminium hydroxide. On the other hand, irreversible colloids cannot be reconverted into the sol condition once they have passed into the gel condition, e.g. stannic hydroxide, gold, etc.

Purple of Cassius.—If a mixture of stannic and stannous chlorides be added to a very dilute solution of gold chloride, hydrated stannic oxide is precipitated and the gold chloride is reduced to the metal. The precipitate of stannic hydroxide may have tints varying from red to violet according to the concentration and composition of the solution. This precipitate is called purple of Cassius because A. Cassius wrote a pamphlet—entitled *De Auro*—describing its preparation in 1685. The mode of

making this substance was known to B. Valentine in 1603; and to J. R. Glauber, 1660. It was used at that time for colouring glass and enamels. The colour of purple of Cassius is due to the precipitation of finely divided gold on the stannic hydroxide. Similar colours are made by precipitating gold on magnesium, calcium, and barium hydroxides, barium sulphate, zirconium oxide, alumina, lead sulphate, and china clay. Purple of Cassius remains as a residue when alloys of gold, and tin with a large excess of silver, are treated with nitric acid; and when a gold-tin alloy is vaporized in air. The tin burns to stannic oxide, and it is at the same time stained by the condensation of the vapour of metallic gold. The old view that purple of Cassius is a compound of tin oxide and aurous oxide has been abandoned. If our opinions were determined by the authority of published papers, about twenty have concluded that purple of Cassius is a compound of gold and an oxide of tin, and seven that it is a mixture of stannic oxide and metallic gold. The twenty chemists just indicated were not able to agree as to the composition of the alleged compound; the analyses range between 24 and 79.4 per cent. of gold; between 20.6 and 76.0 per cent. of stannic oxide; and between 0 and 14.3 per cent. of water. The composition, in fact, varies with the conditions under which it is made. If what are known to be colloidal gold and colloidal stannic oxide are mixed together, the product is the same as that prepared by the usual process. A "purple of Cassius" has been prepared with platinum in place of gold.

§ 9. Copper, Silver and Gold—Oxides and Hydroxides.

The three monoxides Cu_2O , Ag_2O , and Au_2O , are here indicated in the order of their stability. Cuprous oxide, Cu_2O , is the most stable of the copper oxides. It is formed by heating copper or copper oxide in air to a high temperature; silver oxide, Ag_2O , decomposes between 250° and 300° ; and aurous oxide, Au_2O , decomposes at about 250° . Similar remarks apply to the hydroxides. The heats of formation of copper and silver monoxides are: Cu_2O , 40.8 Cals.; Ag_2O , 5.9 Cals. In spite of the fact that silver oxide is decomposed at a red heat, silver appears to unite with oxygen to form Ag_2O at about 1400° , and unless very rapidly cooled, it will decompose at the lower temperature. This phenomenon is connected with the inversion of the thermal value of reactions at elevated temperatures previously discussed. When cupric oxide is reduced by an excess of hypophosphorous acid, the reddish brown sponge-like mass which remains when all the acid is decomposed is supposed to be cupric hydride; the ratio of copper to hydrogen in the analysed product agrees with the empirical formulæ CuH_2 . When washed free from acid and boiled, it changes to a chocolate-brown powder. When heated in air above 100° , it is virtually unchanged until the temperature suffices for the oxidation of the copper. It dissolves in warm concentrated hydrochloric acid with the evolution of hydrogen, the reaction is not so vigorous as the analogous effect with cuprous hydride, and it forms a cupric salt; cuprous hydride forms a cuprous salt. Cupric hydride alone is a moderate reducing agent, but when used in conjunction with zinc or other reducing agents—*e.g.* hypophosphorous acid—it acts as a powerful reducer.

A series of well-defined salts corresponding with Ag_2O are known—silver nitrate, AgNO_3 ; silver sulphate, Ag_2SO_4 , etc. With the exception of silver carbonate, Ag_2CO_3 , these salts do not appear to be hydrolyzed

by water. Cuprous sulphite and cuprous thiosulphate— Cu_2SO_3 and $\text{Cu}_2\text{S}_2\text{O}_3$ —are known; cuprous sulphate, Cu_2SO_4 , probably exists in aqueous solution of cupric sulphate in the presence of metallic copper: $\text{Cu} + \text{CuSO}_4 \rightleftharpoons \text{Cu}_2\text{SO}_4$. This action, however, must be very slight at ordinary temperatures, but in the case of ammoniacal solutions, a colourless crystalline salt, $\text{Cu}_2\text{SO}_4 \cdot 4\text{NH}_3 \cdot \text{H}_2\text{O}$, has been isolated. If methyl (or ethyl) sulphate, $(\text{CH}_3)_2\text{SO}_4$, be brought in contact with cuprous oxide, in the absence of water, at 160° , cuprous sulphate is formed: $(\text{CH}_3)_2\text{SO}_4 + \text{Cu}_2\text{O} = \text{Cu}_2\text{SO}_4 + (\text{CH}_3)_2\text{O}$. The salt is washed with ether, and dried in vacuo. Cuprous sulphate is a greyish white powder, and is fairly stable if moisture be absent. With hydrochloric acid, cuprous chloride, CuCl , is formed; with dilute nitric acid, cupric nitrate; and with sulphuric acid, cupric sulphate; and aurous oxide, Au_2O , like cuprous oxide, exhibits little tendency to form salts. Aurous sulphite and aurous thiosulphate have been reported, but they are probably solutions of colloidal gold reduced by sulphur dioxide or by thiosulphate.

Cuprous oxide is precipitated when an aqueous solution of an alkaline cuprous salt is reduced by glucose or by an arsenite; and it is also precipitated as a yellow powder when an alkaline hydroxide is added to a cuprous salt. Silver salts under the same conditions give a chocolate-brown powder which is somewhat hydrated. It is not silver hydroxide, AgOH ; but is probably argentous oxide, Ag_2O . It is slightly soluble in water (one part of the oxide dissolves in nearly 3000 parts of water); the solution has an alkaline reaction and metallic taste. Silver oxide decomposes when warmed; the action is reversible, $2\text{Ag}_2\text{O} \rightleftharpoons \text{O}_2 + 4\text{Ag}$. There are two solids and one gas, and the system by the phase rule thus resembles the decomposition of calcium carbonate by heat. The pressure of the oxygen in air is about 152 mm. and the dissociation pressure of silver oxide decomposing in a closed vessel at 121° is nearly 152 mm. Hence, silver oxide will decompose completely if heated in air above that temperature, and below that temperature silver will oxidize in air. The rate of oxidation in air is, however, extremely slow unless a catalytic agent be present. Silver oxide is reduced by hydrogen peroxide to metallic silver, and it rapidly oxidizes sulphur, red phosphorus, arsenic sulphide, and antimony sulphide, often with incandescence. Both silver and copper monoxides— Ag_2O and Cu_2O —when precipitated by ammonia, are soluble in an excess of the reagent. On standing, the silver oxide solution deposits black shining crystals of "fulminating silver," which is explosive, particularly when dry. It is often stated to be silver nitride, Ag_3N . Aurous hydroxide, AuOH , is said to be formed when a dilute solution of potassium hydroxide is added to aurous bromide. The violet-brown coloured solution may contain aurous oxide in colloidal suspension. The coagulated powder is decomposed by hydrochloric acid, and is slightly soluble in water and in alkalis, but most of the precipitate appears to be metallic gold. A cold solution of gold in aqua regia is poured into a solution of sodium hydrocarbonate as long as the yellow solution is slightly alkaline. The colour becomes brown on warming, and a light flocculent precipitate separates. This is filtered and washed with water; it is then dried at 110° , and is washed with cold concentrated nitric acid to remove the more soluble gold sesquioxide. The residue, after washing with water and drying, is considered to be a mixture of aurous oxide and gold.

Cupric oxide, CuO , is formed as a black hygroscopic powder when copper is heated in air or in oxygen, or by the calcination of copper nitrate, carbonate, or hydroxide. When heated to a high temperature, cupric oxide cakes, fuses, and decomposes, giving a mixture of cuprous and cupric oxides and finally cuprous oxide. Cupric oxide, when heated, is reduced to the metal by reducing agents—hydrocarbons, carbon monoxide, hydrogen, etc. When an alkaline hydroxide is added to a cold solution of a cupric salt, pale blue cupric hydroxide, $\text{Cu}(\text{OH})_2$, is precipitated. If the liquid be boiled, a black substance is formed, possibly a mixture of $\text{Cu}(\text{OH})_2$ and CuO . It is difficult to wash away the alkalis from the blue precipitate. Boiling water converts it into cupric oxide. The blue precipitate is slightly soluble in an excess of alkali, forming a blue solution. The solution of cupric hydroxide, $\text{Cu}(\text{OH})_2$, in aqueous ammonia is called *Schweitzer's reagent*. It dissolves cellulose (cotton wool, filter paper, etc.). A cellulose-like substance is reprecipitated when the solution is acidified with hydrochloric acid.

Auric oxide, Au_2O_3 .—Silver has been reported to form an oxide corresponding with cupric oxide. When a solution of potassium hydroxide is added to a boiling solution of gold chloride a brown powder—auric oxide, Au_2O_3 —is precipitated. If the solution is cold, an orange-coloured precipitate of auric hydroxide, $\text{Au}(\text{OH})_3$, separates. This is dehydrated at 100° , forming AuO.OH ; at 160° a dark chestnut-brown auro-auric oxide, AuO , is formed; and finally auric oxide, Au_2O_3 . At 250° auric oxide decomposes into its elements. Auric hydroxide is soluble in an excess of alkali, forming a complex salt, potassium aurate, KAuO_2 , which is precipitated by the addition of nitric acid. The aurates are crystalline salts somewhat unstable, and soluble in water. Auric oxide is almost always acidic, and hence it is sometimes called auric acid. Under special conditions it shows feebly basic properties, forming unstable oxy-acid salts like auric nitrate, $\text{Au}(\text{NO}_3)_3$; and auric sulphate, $\text{Au}_2(\text{SO}_4)_3$. The oxide AuO is supposed to be an aurous aurate— Au.AuO_2 . Auric oxide, Au_2O_3 , dissolves in hydrochloric acid, forming a complex acid—hydrochloroauric acid, HAuCl_4 —which forms a series of complex salts, chloroaurates. The chloride AuCl_2 is supposed to be aurous chloroaurate, Au.AuCl_4 . When a solution of gold chloride is treated with ammonia, a yellow precipitate of explosive “fulminating gold” is obtained. Its molecular composition is somewhat uncertain; its empirical composition is said to be $\text{AuN}_2\text{H}_2.3\text{H}_2\text{O}$.

Copper is said to form a copper quadranto-oxide, Cu_4O , according to Rose, when an alkaline solution of cupric chloride is reduced with an alkaline stannous chloride; and silver is said to form silver quadranto-oxide, Ag_4O , when silver citrate is reduced in a current of hydrogen at 100° ; extracted with water; and treated with potassium hydroxide. The evidence is indecisive. Gold, apparently, does not form a similar sub-oxide, although some say that Au_4O exists. The evidence is indecisive, and the alleged suboxides may be mixtures of the respective metals with higher oxides.

Copper dioxide, CuO_2 . In 1844 Krüger attributed the red colour of the solution produced by the action of chlorine gas upon cupric hydroxide suspended in potassium hydroxide to the presence of an unstable copper peroxide. The solution rapidly decomposes, liberating oxygen and precipitating cupric oxide. Cupric dioxide is formed as a yellowish-

brown powder when a dilute solution of hydrogen peroxide acts upon cupric hydroxide. The product prepared in aqueous solution is always impure owing to hydrolysis, and ethered alcoholic solutions are preferred. The compound is obtained by mixing hydrogen dioxide and a solution of cupric chloride cooled to -40° or -50° , and adding alcoholic potassium hydroxide. The precipitate is washed with a mixture of alcohol and ether at -50° . The compound is also made by the action of hydrogen dioxide at a low temperature on copper hydroxide suspended in alcohol, or by the action of an ethereal solution of hydrogen dioxide upon finely divided cupric hydroxide at a low temperature. The compound is unstable, and when treated with dilute acids it gives a cupric salt and hydrogen peroxide. Hence it is probably a superoxide. In addition to the oxides already mentioned, copper has been reported to form Cu_3O and Cu_2O_3 , but the evidence is indecisive.

Silver peroxide, Ag_2O_2 .—The black powder which collects at the anode when silver nitrate is electrolyzed is considered to be silver pernitrate, AgNO_4 , or a mixture of this with silver peroxide. The product soon decomposes, forming a silver peroxide. When dissolved in acids, silver peroxide gives oxygen or ozone; and when slowly acted upon by water it gives hydrogen peroxide: $\text{Ag}_2\text{O}_2 + \text{H}_2\text{O} = \text{Ag}_2\text{O} + \text{H}_2\text{O}_2$. Hence Mendeléeff would consider it to be a superoxide: $\text{Ag}-\text{O}-\text{O}-\text{Ag}$. Silver peroxide dissolves in aqueous ammonia with the evolution of nitrogen: $3\text{Ag}_2\text{O}_2 + 2\text{NH}_3 = 3\text{Ag}_2\text{O} + 3\text{H}_2\text{O} + \text{N}_2$. Gold does not give a peroxide. When potassium persulphate reacts with silver phosphate, argentic oxide with the ultimate composition AgO is formed. This does not give hydrogen peroxide with acids, and it appears to be a basic oxide resembling cupric oxide, CuO . The same oxide seems to be formed when silver is used as anode in the electrolysis of water in an alkaline solution. If the solution be acid, a silver sesquioxide, Ag_2O_3 , is said to be formed.

§ 10. The Family Relationship of Copper, Silver, and Gold.

The properties of the compounds of an element in two different states of oxidation differ absolutely from one another; and in these two states the element appears in two entirely different rôles.—J. LOCKE.

These three elements are generally classed together in so-called "systematic chemistry," and they are furthermore considered to be related with the alkali metals. The copper group of elements is generally linked with the alkali metals by the scheme indicated in the margin, which starts with the elements with the lowest atomic weights. The fault with the scheme is that it makes the relationship appear far closer than is indicated by known facts. If we recall the relationship between the members of, say, the alkali metals and the halogen families, the relationship between the individual members of this so-called family and with the alkali metals is strikingly obscure. The student might well inquire why these elements are grouped together at all, and cogent reasons might be urged for grouping gold with platinum, silver with palladium, and copper with mercury. The reason why this is not done will appear when we take a general survey of all the elements.



Returning to the elements copper, silver, and gold, the physical properties may be tabulated :

TABLE XXVII.—PHYSICAL PROPERTIES OF COPPER, SILVER, AND GOLD.

	Copper.	Silver.	Gold.
Atomic weight	63.6	107.88	197.2
Specific gravity	8.93	10.49	19.265
Atomic volume	7.07	10.29	10.11
Melting point.	1083.0°	958.0°	1063.0°
Boiling point	2310°	1955°	2200°
Latent heat of fusion (cals.) . .	43	24.35	16.3
Specific heat	0.086	0.055	0.030

In malleability, ductility, and tenacity, silver is intermediate between copper and gold. While the large atomic volume of the alkali metals was associated with the great chemical activity and affinity for oxygen, the low atomic volume of these elements is related with their weak affinity for oxygen, etc. Copper, for instance, is alone oxidized in air. The oxides of copper, silver, and gold are easily reduced, while the oxides of the alkali metals are reduced with great difficulty. The reduction of copper, silver, and gold by magnesium is the more energetic the greater the atomic weight of the metal—cuprous oxide reduces easily; silver oxide reduces with explosive violence, and gold oxide breaks down into its constituents below the ignition point of magnesium. Silver appears to be uni-, bi-, and ter-valent, but copper is both uni- and bi-valent, and gold is uni- and ter-valent. Hence these three elements have univalence in common with the alkali metals. Cupric salts are isomorphous with iron, cobalt, and nickel. The isomorphism of the silver and sodium sulphates and selenates indicates a relationship of some kind between silver and sodium. The alkali halides, like silver chloride and bromide, crystallize in the cubic system—silver iodide crystallizes in the hexagonal system and the crystals pass into the cubic system at about 146°—on cooling the reverse change takes place—sometimes with explosive violence. Silver seems to be related with copper through argentic oxide, Ag_2O , where silver appears to be bi-valent; silver pyridine persulphate, $\text{Ag}_2\text{S}_2\text{O}_8 \cdot 4\text{Py}$, is isomorphous with the analogous copper pyridine persulphate, $\text{Cu}_2\text{S}_2\text{O}_8 \cdot 4\text{Py}$; with gold, through silver sesquioxide, Ag_2O_3 ; with the alkalis through argentous oxide, Ag_2O ; and with the magnesium family through mercury. The more salient differences between copper and the alkali metals are : (1) The elements of the alkali metals have a small density, the other metals have a large density ; (2) the alkali metals do not occur free ; (3) the elements of the alkali metals are chemically active, the others not so ; (4) the haloids of the alkalis are all soluble in water, and are not hydrolyzed by water ; copper and silver form sparingly soluble haloids, and the haloids of copper and gold, not silver, are hydrolyzed by water. (5) The oxides and hydroxides of the alkalis are strongly basic ; the oxides and hydroxides of copper and gold are feebly basic, and they accordingly form basic salts. (6) The alkali metals do not form complex salts, whereas copper, silver, and gold form many complex salts.

The remarkable difference between the cuprous and cupric salts—whereby the former make copper resemble silver, and the latter relate it to zinc, ferrous iron, and nickel—emphasizes very forcibly the fact that the properties of an element which forms compounds of a particular valency bear no necessary relation with the properties of the same element which forms compounds with a different valency. This, says G. Martin, can only be because the element, in changing its valency, also “changes the attractive force which it exerts on other atoms, both relatively and absolutely, and this causes a change in those chemical properties which are decided by those forces.” There must be something about an atom of sodium which differentiates it from an atom of silver and this in turn from an atom of chlorine, because under the right conditions the chlorine will leave the sodium and unite with the silver. Now the chemical properties of an atom must be dependent upon the intensity of the force of attraction which the elements have for one another. And these forces must decide which elements will react with a given element, the stability and general character of the resulting compounds, etc. G. Martin (1905) has shown that the more nearly alike the attractive forces of two elements A and B for a third element C, the more nearly will the compounds AC and BC resemble one another, and that the chemical similarity of elements must be determined not by their valency nor their atomic weights, but rather on their capacity for exerting an attractive force of the same or of proportional intensity. Assuming that the attractive forces are proportional to their heats of formation, Martin shows that the ratio of the heats of formation of the cuprous and silver halides are:

$$\frac{[\text{Cu}, \text{Cl}]}{[\text{Ag}, \text{Cl}]} = \frac{32.9}{29.2} = 1.12; \quad \frac{[\text{Cu}, \text{Br}]}{[\text{Ag}, \text{Br}]} = \frac{25.0}{23.0} = 1.10$$

This is taken to mean that the attractive forces of the silver atom on the atom of chlorine and bromine are roughly proportional to those of the cuprous copper atom on the same elementary atom, and hence the elements are chemically similar. Comparing ferrous iron and cupric compound in a similar way,

$$\frac{[\text{Cu}, \text{O}]}{[\text{Fe}, \text{O}]} = \frac{37.2}{64.6} = 0.58; \quad \frac{[\text{Cu}, \text{Cl}_2]}{[\text{Fe}, \text{Cl}_2]} = \frac{51.6}{82.0} = 0.63; \quad \frac{[\text{Cu}, \text{Se}]}{[\text{Fe}, \text{Se}]} = \frac{9.7}{19.4} = 0.52$$

which shows, says Martin, that the cupric atoms attract other atoms with a force which is in proportion to that exerted by ferrous atoms, but only half as great. This shows how cupric atoms have similar properties to the ferrous atoms, and different properties from the univalent silver atoms.

It is curious to trace the kinship of monad mercury with silver. The halides are remarkably similar. The two chlorides are white, the bromides distinctly yellow, and the iodides yellow. The yellow mercurous halides are more intensely coloured than the corresponding silver halides. The same remark applies to the nitrites and hyponitrites. The mode of formation of these salts are similar. Silver and mercurous nitrites (not the nitrates) appear to be the first products of the action of cold nitric acid on the metals. The chemical and physical properties of these salts are also analogous. In some complex mercurous, or mercuric oxynitrates, silver can isomorphously replace monad mercury. Silver and monad copper can isomorphously replace one another in several salts, *e.g.* in some complex

thiocyanates and thiosulphates. Monad mercury is closely related to silver, far more closely, indeed, than gold to silver. Hence P. C. Rây (1914) has argued that gold should be displaced from this family and relegated to the platinum group, and that the natural place for monad mercury is alongside silver.

Questions.

1. Describe the successive phenomena observed when ammonium hydroxide solution is added, drop by drop, to a solution of cupric sulphate, and indicate by formulae or equations the nature of the changes which occur.—*Massachusetts Polytechnic Inst., U.S.A.*

2.—(a) In the laboratory you passed dry hydrogen over copper oxide which was heated in a tube. How does this experiment illustrate an analytical reaction; a synthetic reaction; oxidation and reduction? (b) If you started with an unknown mixture of copper oxide and copper, say 10 grams, and after heating and passing hydrogen over it, the resulting weight of pure copper was 9.2 grams, how much of the original weight of the 10 grams was copper oxide and how much was metallic copper?—*Worcester Polytechnic, U.S.A.*

3.—(a) Show the analogy between the reactions of copper on nitric acid and copper on sulphuric acid. (b) What products are formed from the concentrated sulphuric acid when the latter is warmed with potassium iodide? (c) What reason have you for thinking that no nascent hydrogen is formed in the reactions mentioned under (a)?—*Amherst Coll., U.S.A.*

4. It is desired to obtain pure copper nitrate, the only source of copper at hand being a complex solution of lead, cadmium, copper, and mercurous nitrates. Using this solution as the source of the copper, how may a new solution of copper nitrate be prepared?—*Massachusetts Inst. of Technology, U.S.A.*

5. An unknown quantity of potassium bromoaurate, KAuBr_4 , on being heated, left 9.9245 grams of a mixture of metallic gold and potassium bromide. The mass on being treated with water left 6.18997 grams of gold. The solution of KBr required 3.38540 grams of silver for total precipitation by Stas' method, and afforded 5.89143 grams of silver bromide. These data afford three independent values for the atomic weight of gold, which you are required to calculate ($\text{K} = 39.03$ $\text{H} = 1$; $\text{O} = 15.96$; $\text{Br} = 79.76$; $\text{Ag} = 107.66$).—*Science and Art Dept.*

6. What takes place when hydrogen is passed over red hot copper oxide? How has this reaction been applied to determine (a) the composition of water, (b) the atomic weight of copper? What special precautions must be observed in each case in order to obtain accurate results?—*London Univ.*

7. A chemical manufacturer gives out a kilogram of metallic silver to be made into silver chloride: how much silver chloride ought to be returned by the workman, supposing the average yield is 99.8 per cent. of the theoretical?

CHAPTER XXII

SULPHUR AND ITS HYDROGEN COMPOUNDS

Atomic weight, $S = 32.07$; molecular weight of solid, $S_8 = 256.56$; of vapour at 1000° , $S_2 = 64.14$. Melting point of rhombic sulphur, 112.8° ; monoclinic sulphur, 119.2° ; boiling point, 444.7° . Relative vapour density depends on temperature (air = 1), 6.6 to 2.2.

§ 1. Sulphur—Occurrence and Preparation.

THE element sulphur has been known from the beginning of history. It is mentioned in the Bible and in Homer. It was placed among the elements by Lavoisier, but for some time previously it was regarded as "the principle of fire." The name is derived from the Sanscrit *sulveri* through the Latin *sulphurium*.

Occurrence.—Sulphur is widely distributed in nature both as free and as combined sulphur. Deposits of free or native sulphur occur in volcanic

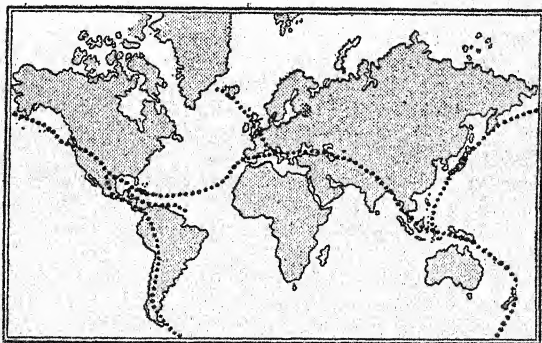


FIG. 160.—Imaginary Lines showing Chief Regions of Volcanic Phenomena and Occurrence of Sulphur Deposits.

districts, Iceland, Italy (Romagna, Marken, Tuscany, Campania, and Calabria), Sicily (chiefly on the southern watershed), Greece (Island of Milo), Russia, Austria-Hungary (Radoboj and Swoscowice), South France, Spain, Asia Minor, Persia, India, Palestine, Algeria, Morocco, Japan (Sulphur Island), New Zealand (White Island, etc.), United States (Louisiana, Oregon, Utah, Nevada, Wyoming, Texas), Mexico (Popocatepetl), Chili, Peru, etc. These districts are on or not far from the dotted line, Fig. 160.

There are two main types of native sulphur: (1) *The solfataric type*¹ found in lava fissures and in extinct volcanic vents (Japan, Mexico). Deposits of this type are forming at the present day in volcanic districts from the mutual action of hydrogen sulphide and sulphur dioxide which occur among volcanic gases. The water from a sulphur spring at Kriswoik, Iceland, is said to be ejected with a hissing noise to a height of 5–8 ft. and to be accompanied by steam impregnated with sulphur dioxide and hydrogen sulphide. The water subsequently deposits sulphur, and in many places the surface of the ground has a crust of sulphur 2–3 ft. thick. This was at one time collected and exported. (2) *The gypsum type*. The sulphur appears to have been liberated from gypsum by the reducing action of bituminous matters found associated with the gypsum. Calcium sulphide is probably formed from the sulphate; and this, by the action of water and carbon dioxide, forms sulphur, calcium carbonate and hydrogen sulphide. The Sicilian deposits and perhaps the more important sulphur deposits are of this type.

Sulphur is also a product of the life action of certain bacteria—*Beggiatoa alba* and *Chromatium Okeini*. Just as plants and animals derive energy and heat by the oxidation of carbon, so do these bacteria appear to get their energy and heat by the oxidation of sulphur. It is estimated that 25 per cent. of their body is sulphur. These bacteria can oxidize hydrogen sulphide to sulphuric acid, and an intermediate stage of the oxidation process is the change of hydrogen sulphide to sulphur which is stored inside the cells. The members of another variety of sulphur bacteria—the so-called *thiobacteria*—do not store sulphur internally; but they oxidize thiosulphates to sulphuric acid. The bacteria theory of the origin of sulphur assumes that some sulphur beds have been formed by these organisms.

Many important metallic ores are sulphides, that is, compounds of sulphur with one or more metals. Thus, *galena* (lead sulphide, PbS); *zinc blende* (zinc sulphide, ZnS); *cinnabar* (mercury sulphide, HgS); *stibnite* (antimony sulphide, Sb_2S_3); *copper pyrites* (Cu_2S , Fe_2S_3); *iron pyrites* (iron sulphide, FeS_2), etc. Some important sulphates also contain sulphur, e.g. *gypsum* (calcium sulphate, CaSO_4); *heavy spar* (barium sulphate, BaSO_4), etc.

Sulphur occurs in many organic compounds, and in animal and vegetable products: onions, garlic, mustard, horseradish, hair, many oils, eggs, proteids, etc. Hydrogen sulphide is found in the water of many sulphur springs.

The extraction of crude sulphur.—A little sulphur is obtained from iron pyrites; from the by-products in Leblanc's soda process (*q.v.*); and from the spent oxide of gasworks (*q.v.*). Probably 90 per cent. of the sulphur used in the world comes from Sicily and Louisiana. The "sulphur earth" in Sicily occurs in lodes mixed with limestone and gypsum. The amount of sulphur in "workable" ore varies from 8 up to about 25 per cent. The sulphur is separated by heating the ore, and allowing the molten sulphur to flow away from the mineral impurities. This is effected by stacking the ore on the sloping floor, Fig. 161, of a circular kiln without a permanent roof. The kilns are called *calcaroni*. In stacking the ore, air spaces are left at intervals to serve for ventilation. The stack is covered with powdered or burnt ore. The sulphur is ignited near the bottom. A

¹ Solfatara—a volcanic vent or volcanic area which gives off sulphurous vapours, steam, etc., and which probably represents the last stages of volcanic activity

portion of the sulphur acts as fuel, and melts the remainder which collects at the lowest point of the inclined bottom of the kiln. After about five days, a plug at the lower end of the kiln is removed, and the sulphur is run into small wooden moulds. The opening is closed to be reopened day by day until, in from three to five more days, the sulphur ceases to flow.

About one-third of the sulphur is lost in the calcarone system of extraction. It is, however, cheaper to use the sulphur as a fuel than to import coal. The loss, however, is excessive even then, and in consequence, the calcarone method is being displaced by more economical kilns—Gill's kilns—which are worked in sets. The products of combustion from one kiln pass into the adjacent kiln and there do some work before escaping

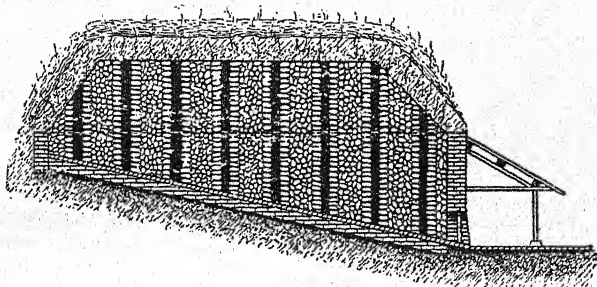


FIG. 161.—Calcarone or Sulphur Kiln (Diagrammatic Section).

into the atmosphere. In H. Frasch's method (1891), used at Louisiana, the sulphur is melted *in situ*. Pipes are driven into the sand and superheated water is forced into the lode. The molten sulphur collects in a central well from which it is raised to the surface by compressed air.

The purification or refining of crude sulphur.—The crude sulphur from the kilns—also called “brimstone”—is graded and put on the market. It may be afterwards purified by distillation from a retort which opens into a large brickwork chamber. The sulphur vapour condenses in the chamber. The first lot of vapour sublimates as a light powder on the walls. The powder is called *flowers of sulphur*. As the condensing chamber gets hot, the condensed sulphur melts and collects on the floor as a liquid which is drawn off from time to time and cast into large wooden moulds—*rock sulphur*; or in cylindrical wooden moulds—*roll sulphur*.

Uses.—Crude sulphur is used in making sulphur dioxide for bleaching straw, wool, etc.; for the manufacture of sulphites for bleaching wood fibres, etc.; and for the manufacture of sulphuric acid. It is also used in making carbon disulphide. Purified sulphur is used in making gunpowder, matches, colours, vulcanite, etc.; and also medicinally. Flowers of sulphur is used as an insecticide and fungicide.

§ 2. Rhombic, Octahedral, or α -Sulphur.

Ordinary sulphur is a pale-yellow brittle solid, without taste or smell. At -50° the sulphur is almost colourless. Sulphur is commonly found in rhombic or octahedral crystals with a specific gravity varying from

2.03 to 2.06. It is a bad conductor of heat and electricity. If a stick of roll sulphur be held in the hand it begins to crackle and breaks owing to unequal expansion by the heat. Sulphur is practically insoluble in water; sparingly soluble in alcohol and glycerine; more soluble in essential oils; and readily soluble in sulphur chloride, S_2Cl_2 , benzoyl chloride, and carbon disulphide, CS_2 . For instance, 100 grams of carbon disulphide at 0° dissolve 22 grams of sulphur; at 20° , 41.8 grams; and at 40° , 100 grams. Benzoyl chloride dissolves one per cent. of sulphur at 0° , and 55.8 per cent. at 134° ; above 136° , solvent and solute are miscible in all proportions. If the solution in chloroform, benzoyl chloride, or carbon disulphide be allowed to stand at ordinary temperatures, crystals of sulphur are deposited as the solvent evaporates. The appearance of the crystals is shown by the

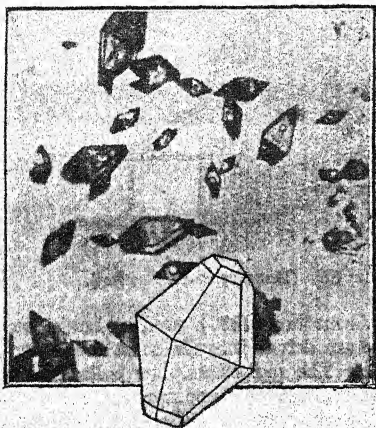


FIG. 162.—Rhombic or α -sulphur.

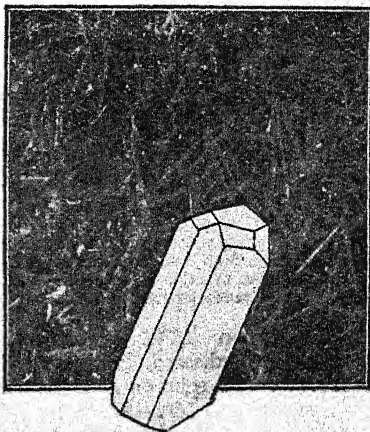
photograph, Fig. 162. An ideal crystal is sketched in outline. About 60 or 70 per cent. of flowers of sulphur consists of these crystals. Native sulphur also occurs in this form. The crystals are quite stable at ordinary temperatures. This crystalline form of sulphur is called " α -sulphur," "rhombic sulphur," or "octahedral sulphur." Nearly thirty different crystalline modifications of the rhombic type of crystalline sulphur have been reported, and this fact leaves us some misgivings about the rigid exactness of R. J. Haüy's law (p. 206).

Sulphur unites with oxygen when heated in air. If thoroughly purified and dried, sul-

phur may be distilled, without oxidation, in thoroughly dried oxygen. Sulphur in air ignites at about 363° , and at about 282° if heated in oxygen. The vapour of sulphur ignites at about 285° in air. Too much reliance should not be placed on these numbers since they are largely dependent upon the conditions of the experiment, so that different experimenters report widely different numbers. At 100° sulphur is oxidized in oxygen gas at a measurable rate. Oxidation, even at ordinary temperatures, can be detected. Finely divided sulphur oxidizes slowly in moist air, forming sulphurous and sulphuric acids. Hence these two acids can nearly always be detected in commercial flowers of sulphur, while roll sulphur is almost free from these acids. Sulphur also unites readily with many metals, forming sulphides. Heat is usually required to start the reaction. The combination is often attended with incandescence. Examples with iron and zinc have been indicated on p. 33. A strip of copper introduced into the vapour of sulphur enters into combination with vivid combustion. Sulphur combines with carbon at a red heat, forming carbon disulphide— CS_2 ; with chlorine at the boiling point of sulphur; and with hydrogen at the same temperature.

§ 3. Monoclinic, Prismatic, or β -Sulphur.

In 1823 E. Mitscherlich announced the fact that the element sulphur can be crystallized in two distinct forms; and concluded that a substance, whether simple or compound, may assume two different crystalline forms, p. 214. If, say, 500 grams of sulphur be melted in a clay or porcelain crucible and the mass allowed to stand until a surface crust is formed, beautiful long prismatic needle-like crystals of waxy yellow sulphur will be found to have grown on the walls of the crucible, and on the underside of the crust when the crust is pierced, and the still fluid sulphur is poured away. These monoclinic prisms of sulphur have many properties different from ordinary rhombic sulphur just discussed. The specific gravity, for instance, is 1.93, instead of 2.04; and the melting point is 109.2° in place of 112.8° . Rhombic sulphur is soluble in carbon disulphide, and the monoclinic form is rapidly changed by this solvent to the rhombic form when it dissolves. The solubility of monoclinic sulphur in benzene, chloroform, ether, and alcohol is about 1–3 times greater than that of rhombic sulphur—e.g. in benzene at about 18° , 10 c.c. of solution contains 0.2004 gram of β -sulphur or 0.1512 gram of α -sulphur; and 10 c.c. of the alcoholic solution at 25° contains 0.0066 gram of β - and 0.0052 gram of α -sulphur. Some crystals of monoclinic sulphur are illustrated in Fig. 163 along with an outline sketch of an ideally perfect crystal. This form of sulphur is called " β -sulphur," "monoclinic sulphur," or "prismatic sulphur."

FIG. 163.—Monoclinic or β -sulphur.

In about a day's time, the monoclinic prisms become light yellow, opaque, brittle, and crumble into powder at the slightest touch. The grains of powder are small rhombic crystals of α -sulphur. If the rhombic crystals be kept a few hours between 108° and 112° , they also become opaque and change to a friable crumbling mass of monoclinic prisms. The two reactions are thus reversible. Experiment shows that the monoclinic prisms are unstable below 94.5° , and slowly pass into the rhombic variety. The change is accelerated by wetting the monoclinic prisms with carbon disulphide, or by bringing the monoclinic sulphur in contact with a crystal of the rhombic variety. Conversely, the rhombic sulphur is unstable above 94.5° , and slowly passes into the monoclinic variety. With the notation previously employed, the change is symbolized:



Hence, 94.5° is a transition temperature.

If a little sulphur be melted on a glass slide and allowed to cool gradually, monoclinic crystals, Fig. 163, first appear; as the temperature falls, these change to rhombic crystals. It is possible to regulate the temperature so that a certain amount of each form exists simultaneously, but on raising the temperature, the monoclinic crystals grow at the expense of the rhombic crystals, and conversely on lowering the temperature.

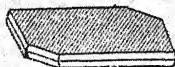


Fig. 164.—Nacreous Sulphur.

The monoclinic crystals just indicated were studied by E. Mitscherlich in 1823. There are, however, two other varieties of monoclinic crystals with angles somewhat different from the variety just indicated. In 1884 D. Gernez prepared what he called "soufre nacré," nacreous sulphur—from the French *nacré*, mother-of-pearl—as follows:—

Heat sulphur in a sealed tube with benzene, or toluene, carbon disulphide, alcohol, etc., so that there is no undissolved sulphur in the tube when the tube is hot. Then immerse one end of the tube in a freezing mixture formed, say, by dissolving ammonium nitrate in cold water. Long nacreous flakes separate at the cold end of the tube and gradually extend into the remaining solution.

An ideal monoclinic crystal of nacreous sulphur is illustrated in Fig. 164. Similar crystals were made by E. Mitscherlich in 1823 by the following method:—

Saturate an alcoholic solution of sodium sulphide with sulphur; filter off the clear reddish coloured supernatant liquid, and, after adding a little more alcohol, let the solution stand for some time. Needle-like crystals of nacreous sulphur grow from the surface of the solution.

W. Muthmann (1890) noticed that in preparing nacreous sulphur by the method just indicated, hexagonal plates—tabular sulphur—sometimes separate from the solution when the temperature is about 5°, but not above 14°. These crystals also belong to the monoclinic system. An ideal crystal is illustrated in Fig. 165. Although the three types of crystals just named belong to one system, the interfacial angles are not the same. Thus, the angle between two important faces of one of the crystals depicted in Fig. 163 is 84° 14'; with similar faces of the nacreous crystals, Fig. 164, 88° 13'; and with tabular crystals, Fig. 165, 75° 58'.



Fig. 165.—Tabular Sulphur.

The rhombic crystals (Fig. 162) are sometimes called Muthmann's S_I ; the monoclinic crystals (Fig. 163), Muthmann's S_{II} ; the nacreous crystals (Fig. 164), Muthmann's S_{III} ; and the tabular crystals (Fig. 165), Muthmann's S_{IV} . In addition to the crystalline varieties of sulphur just considered, two others, of no particular importance in our present study, have been reported, namely, G. Friedel's *triclinic sulphur* (1879), and R. Engel's *rhombohedral sulphur* (1891). When an element or compound exists in two or more crystalline forms it is said to be polymorphous—from the Greek *πολύς* (polus), many; and *μορφή* (morphē), form. Hence sulphur is polymorphous. If a substance exists in two different crystalline forms, it is said to be dimorphous—from the Greek *δύς* (dis), twice. Ammonium nitrate and calcium carbonate are examples.

§ 4. Sulphur and the Phase Rule.

By plotting the vapour pressure curve of rhombic sulphur at different temperatures, we get the curve PO_1 , Fig. 166; similarly, by plotting the vapour pressure curve of monoclinic sulphur, we get the curve O_1O_2 ; this variety of sulphur melts at 120° ; however, by continuing the vapour pressure curve of the liquid, we get O_2Q . By plotting the transition points of rhombic sulphur at different pressures, we get the curve O_1O_3 ; and by plotting the melting point of monoclinic sulphur at different pressures, we get the curve O_2O_3 . Monoclinic sulphur cannot exist at pressures higher than that represented by the point O_3 . The continuation of the curve O_3N represents the effect of pressure on the melting point of rhombic sulphur. This diagram, Fig. 166, should be compared with Fig. 61 for water. In Fig. 166, we have the additional complication corresponding with the two forms of sulphur now under consideration. The phase rule enables us to form a very clear idea of the conditions of equilibrium. When the condition of the system is represented by a pressure and temperature corresponding with one of the three triple points— O_1 , O_2 , O_3 —the system is invariant, and any change in temperature or pressure will lead to the suppression of one of the three phases; points on one of the curves PO_1 , O_1O_2 , O_2Q , etc.—represent univariant systems; and points in one of the three regions— PQ , QO_2N , NO_1P —represent bivariant systems. It will of course be obvious that we are here dealing with the one component sulphur, and four possible phases

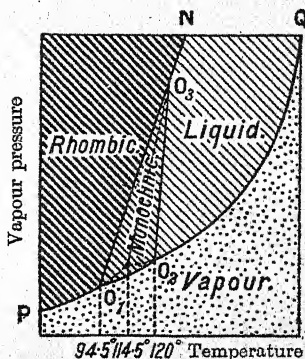


Fig. 166.—Vapour Pressure of Sulphur (Diagrammatic).

—sulphur vapour and liquid, and rhombic and monoclinic sulphur—since an allotropic modification of an element is a true phase. Can all four phases exist under any conditions of temperature and pressure in a state of equilibrium? According to the phase rule, the variance of such a system will be $1 - 4 + 2 = -1$. This is an impossible value. Such a system would not be in a state of true equilibrium. The metastable states, or states of false equilibrium, are interesting. The QO_2 and the PO_1 curves meet at a point corresponding with the temperature 114.5° . This is the melting point of rhombic sulphur. If the transformation of rhombic to monoclinic were very fast, it would be impossible to state the melting point of rhombic sulphur, because it would pass into the monoclinic form before a determination could be made.

The upward left-to-right slopes of the curves O_1O_3 and O_2O_3 ¹ correspond with the fact that the melting point of sulphur is raised by increasing pressures. The converse was true in the case of ice, Fig. 61, and in consequence, the corresponding curve sloped the opposite way. In illustration

¹ These curves are exaggerated in the diagram. That, however, does not affect the principles under discussion.

of the principle of least action, viz. an increase of pressure favours the progress of that change which occurs with a decrease in volume. W. Spring found that monoclinic sulphur is converted by pressure into the denser rhombic form, for the decrease in volume which occurs during the transformation tends to counteract the disturbing influence of pressure. Water, too, expands when it passes into ice, so that pressure changes ice into denser water. Pressure also changes black arsenic into denser grey arsenic. Again, if a compound occupies a less volume than the joint volume of its constituents, the separate constituents may unite under pressure, and conversely. Hydrated arsenious sulphide, $\text{As}_2\text{S}_3 \cdot 6\text{H}_2\text{O}$, is decomposed by pressure into water and arsenious sulphide. Powdered, dry alumina, silica, or ferric oxide do not unite under pressure, but they do so if wet. A solution usually occupies a smaller volume than the joint volume of solvent and solute taken separately. Hence solids usually dissolve with a contraction. When a wet powder is compressed, the surface layers of the grains dissolve; when the pressure is relieved, the solute may be rejected and form a kind of cement binding the mass together. Wet powders of potassium iodide, or ammonium chloride, do not cohere. In these cases the solids dissolve with an expansion and pressure which do not favour the formation of a solution.

§ 5. Amorphous or Colloidal Sulphur.

Sulphur exists in at least two different colloidal or amorphous conditions in which no signs of crystallization can be detected under the microscope. The term amorphous is derived from the Greek *α*, without; *μορφή* (morphē), form. One of the amorphous forms is soluble in carbon disulphide, the other is not. If hydrogen sulphide be passed into a saturated aqueous solution of sulphur dioxide at 0° ; or if an alcoholic solution of sulphur be poured into water; or if a saturated solution of sodium thiosulphate be mixed with twice its volume of concentrated hydrochloric acid, and cooled to 10° , colloidal, soluble, or δ -sulphur is formed.

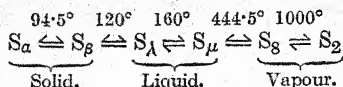
Action of heat.—Sulphur is pale yellow at ordinary temperatures, and almost colourless at -50° , and at 100° it is intense yellow. If a piece of ordinary rhombic sulphur be gradually heated in a test-tube, the sulphur crackles and falls to pieces as indicated above. As the temperature rises, the sulphur melts to a clear, limpid, amber-coloured liquid between 113° and 115° ; the colour darkens, and the liquid loses its mobility, until, at about 160° , the mass is almost black, and so viscous that the test-tube can be turned upside down without pouring out the sulphur. The viscosity reaches a maximum at about 180° , for as the temperature rises still higher, the dark colour remains, but the mass becomes more and more mobile until at 444.7° the liquid begins to boil, forming a reddish orange vapour. If the liquid be allowed to cool, the sulphur undergoes the same changes, but in the reverse order, and very slowly. If the vapour be heated still higher, it becomes deep red at 500° , and straw-yellow at about 650° . Sulphur viewed in darkness appears phosphorescent when its temperature reaches 200° , and ozone is formed at the same time. No signs of a lower oxide of sulphur, nor of an ionization (i.e. increased electrical conductivity) of the atmosphere is observed during the low temperature oxidation.

Molten sulphur— λ - and μ -sulphur.—If sulphur, heated to about 350° , be poured into cold water, a tough elastic material resembling indiarubber—called plastic sulphur—is obtained. Plastic sulphur is also obtained by distilling ordinary sulphur from a glass retort and allowing the liquid sulphur, condensing in the neck of the retort, to flow into cold water. A long continuous thread of plastic sulphur is then obtained. The specific gravity of plastic sulphur is about 1.95, nearly the same as monoclinic sulphur; but unlike the crystalline varieties, this form of sulphur can be moulded between the fingers, and drawn into somewhat elastic threads. Plastic sulphur is a supercooled liquid which has been hurried past its crystallizing temperature and cooled so low that it has formed a viscid mass. If cooled slowly, virtually the whole of the product is crystalline and soluble in carbon disulphide. The case is analogous with that studied on p. 198—supercooled sodium thiosulphate. Plastic sulphur slowly¹ crystallizes on standing. The change is accelerated by rubbing the mass, and is fairly rapid if the mass be heated to about 100° . The latent heat of solidification—corresponding with about 9.4 Cals. per kilogram—is evolved when the plastic or supercooled sulphur crystallizes. Amorphous “solids” are in nearly every case to be regarded as supercooled liquids which have not taken up the stable crystalline condition.

If α -sulphur be melted at a low temperature, and the pale yellow liquid be suddenly chilled, the crystalline product is almost completely soluble in carbon disulphide; while, as indicated above, if the dark brown liquid which is obtained when sulphur is heated to a higher temperature be similarly treated, it forms an amorphous mass almost all insoluble in carbon disulphide. It is therefore inferred that molten sulphur contains a mixture of two varieties of sulphur—the pale yellow mobile fluid, called λ -sulphur, and the dark brown viscid fluid called μ -sulphur. When λ -sulphur solidifies it forms crystalline α - or β -sulphur soluble in carbon disulphide; and when μ -sulphur solidifies it forms an amorphous plastic mass insoluble in carbon disulphide. The proportion of sulphide soluble and insoluble in carbon disulphide in solidified sulphur depends on the relative amounts of S_λ and S_μ present in the fluid at the time of solidification. The proportion of the two varieties in any sample of sulphur can be determined by extraction with carbon disulphide, making a small allowance for the very slight solubility of the so-called “insoluble” variety. Experiment shows that there is a definite relation between the relative amounts of the two varieties present in a system in equilibrium at a definite temperature. Thus, molten sulphur at 114.5° , if suddenly congealed, furnishes a mass which contains the equivalent of 3.7 per cent. of μ -sulphur, and 94.3 per cent. of λ -sulphur. Ordinary plastic sulphur contains the equivalent of about one per cent. of λ -sulphur and 99 per cent. of μ -sulphur. Ordinary rhombic sulphur has about 3.4 per cent. of the first variety, and it is estimated that if the latter were absent the melting point would be 119.25° , not 114.5° ; similarly with monoclinic sulphur, the melting point would be raised a couple of degrees. The proportion of μ -sulphur appears to increase with rising temperatures, and for each temperature there appears to be a definite equilibrium constant corresponding with the reversible reaction: $S_\lambda \rightleftharpoons S_\mu$. The system takes some time to attain equilibrium under any given conditions. The presence of air, sulphur

¹ A trace of iodine retards the action.

dioxide, or hydrogen chloride retards, while ammonia, nitrogen, carbon dioxide, iodine, or hydrogen sulphide accelerates the speed of the change. We may thus summarize the action of heat on the different forms of sulphur:



The μ -variety of sulphur is only sparingly soluble in sulphur chloride, and it separates in crystals when a hot solution is cooled, but if a solution of sulphur in sulphur chloride, saturated at ordinary temperature, be heated to 150° and then cooled, it dissolves much more sulphur. A. H. W. Aten (1912) considers this is due to the formation of another modification of sulphur which he calls π -sulphur. Sulphur which has been heated to 170° and rapidly cooled becomes more soluble in sulphur chloride and in toluene. This is supposed to be due to the formation of π -sulphur which in turn is believed to be identical with the "crumbly sulphur" described by Magnus (1856). Solutions of this variety of sulphur are deep yellow (like potassium chromate solutions). When a carbon disulphide solution of this sulphur is evaporated in vacuo at -80° the residual sulphur is almost wholly soluble in toluene, the insoluble residue is μ -sulphur. Solutions of λ -sulphur and of mixed λ - and π -sulphur on exposure to light precipitate μ -sulphur, but less rapidly in the latter than in the first-named solution.

Amorphous sulphur soluble and insoluble in carbon disulphide.—Let slaked lime be boiled with water and sulphur for some time. After

TABLE XXVIII.—VARIETIES OF SULPHUR. (To face page 481.)

	Variety.	Carbon disulphide.	Water.	Sp. gr.	Synonyms.
Crystalline.	Rhombic . . .	Soluble	Insol.	2.04	Octahedral S; α -S; Muthmann's S_I
	Monoclinic needles	Soluble	Insol.	1.93	Prismatic S; β -S; Muthmann's S_{II}
	Monoclinic nacreous Monoclinic tabular	Soluble Soluble	Insol. Insol.	— —	Nacreous S; Muthmann's S_{III} Tabular S; Muthmann's S_V
Colloidal.	Amorphous . . .	Soluble	Insol.	—	γ -sulphur δ -sulphur
	Amorphous . . .	Insol.	Insol.	1.955	
	Soluble . . .	Soluble	Soluble	—	
Fluid	Fluid λ -S	Solidifies to soluble amorphous or crystalline S
	Fluid μ -S	Solidifies to insoluble amorphous S

the mixture has stood to allow the insoluble matters to settle, decant the clear reddish solution of calcium polysulphide. Add a little hydrochloric acid. Calcium chloride is formed, and finely divided *milk of sulphur* is precipitated. The sulphur remains suspended in the liquid for a long time on account of its very fine state of subdivision. This sulphur is

pale yellow in colour, it is said to contain a non-crystalline variety of sulphur soluble in carbon disulphide. A certain amount of the variety (sp. gr. 1.955) insoluble in carbon disulphide is formed at the same time; this has been called γ -sulphur. Several other sub-varieties of amorphous sulphur are known. These are more or less closely related to the leading types just considered. A complete and satisfactory statement of the different varieties of crystalline and amorphous sulphur is not yet possible. Those indicated above are tabulated on p. 480 (Table XXVIII.).

§ 6. The Atomic and Molecular Weights of Sulphur.

1. Combining weight.—J. B. Dumas (1859) and J. S. Stas (1867) heated silver in a tube through which a current of sulphur vapour was passed. The excess of sulphur was distilled off in a current of carbon disulphide. The resulting silver sulphide was weighed. It was found that 100 grams of silver gave 114.85 grams of silver sulphide. Hence $\text{Ag} : \text{S} = 100 : 14.85$. The combining weight of silver (oxygen = 8) is 107.88; hence the combining weight of sulphur is 16.035. Results not very different have been obtained by reducing a known amount of silver sulphide in a current of hydrogen; by finding the amount of silver in silver sulphate; by converting silver sulphide into silver chloride; and by converting sodium carbonate into sodium sulphate (T. W. Richards, 1891).

2. Atomic weight.—Applying the method indicated on p. 80, namely, collecting together a number of volatile compounds of sulphur whose vapour density is known, we obtain :

TABLE XXIX.—VAPOUR DENSITIES OF SOME VOLATILE SULPHUR COMPOUNDS.

Volatile compound.	Vapour density, $H = 2$.	Composition : molecular weight = vapour density.	Amount of sulphur in the molecule.
Hydrogen sulphide .	34.07	Sulphur 32.07; hydrogen 2.016	32.07
Sulphur dioxide . .	64.07	Sulphur 32.07; oxygen 32.0	32.07
Sulphur trioxide . .	80.07	Sulphur 32.07; oxygen 48.0	32.07
Sulphur	64.14	Sulphur 64.14	64.14
Carbon disulphide .	76.14	Sulphur 64.14; carbon 12.0	64.14
Phosphoric sulphide	222.35	Sulphur 160.35; phosphorus 62.0	160.35

The smallest amount of sulphur entering into the composition of any of these molecules lies somewhere between 32.01 and 32.14; the best representative value is taken to be 32.07 when hydrogen is 1.008; oxygen, 16; carbon, 12.0; phosphorus, 31.0. Hence this number represents the atomic weight of sulphur. This result is confirmed by accurate determinations of the density of sulphur dioxide. This is 64.07. Hence if oxygen be 16, sulphur must be 32.07.

3. Molecular weight.—At about 500° the vapour density of sulphur is nearly 6.6 (air = 1). This corresponds with the molecule S_8 . By raising the temperature or reducing the pressure, the vapour density gradually diminishes until, at 1000° , it reaches 2.2, corresponding with the molecule S_2 . The vapour density then remains constant, up to about

1700°; at about 2000° the vapour density corresponds with a partial dissociation of 50 per cent. of the S_2 into atoms S . The phenomenon with sulphur thus resembles the behaviour of iodine above 1000°. It may be that at temperatures intermediate between 1000° and 500°, some or all of

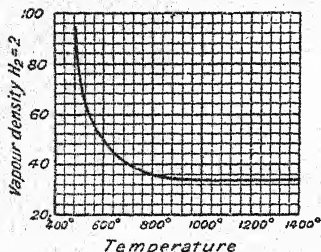
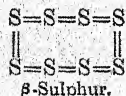


Fig. 167.—Vapour Density of Sulphur.

the molecules S_8 , S_6 , S_4 , and S_2 are present, but the curve is steadily depressed with rising temperatures, and it shows no signs of flattening such as might be expected if any particular type of molecule predominated throughout any particular range of temperature. This is illustrated by the graph of the vapour density of sulphur at different temperatures shown in Fig. 167. The freezing and boiling point methods for the determination of molecular weights, indicated by examples pp. 259 and 260, show that the molecule of sulphur in a solution of carbon disulphide, bromoform, etc., exists as S_8 . H. Erdmann (1908) represents the molecules of the two different forms of crystalline sulphur graphically by the formulæ:



The former is supposed to represent monoclinic sulphur, and the latter rhombic sulphur. The experimental evidence upon which these formulæ are based is very flimsy. The molecules of both forms by the freezing and boiling point methods give the same results: S_8 .

§ 7. Allotropy.

It matters much with what others and in what positions the same atoms are held together. When the configurations of the atoms of matter are changed, the things which are formed from them must also change.—LUCRETIVS.

The relation between ozone and oxygen and between the different forms of sulphur must be interesting. Ozone and oxygen are dual forms or dual personalities of one elemental form of matter. This was proved on p. 225. The different modifications of sulphur likewise contain but one elemental form of matter. This can be proved by showing that a known weight of any of the different forms of sulphur furnish the same amount of sulphur dioxide when burnt in oxygen gas. The experiment can be made by an apparatus resembling that depicted in Fig. 168. About 0.1 gram of pure dry sulphur is introduced into a porcelain boat, and all is weighed. The boat is introduced into a hard glass tube, *C*, which is connected at one end with a gas holder, *A*, containing oxygen, and wash-bottles, *B*, containing concentrated sulphuric acid to dry the oxygen gas. The other end of the combustion tube is connected with weighed glass bulbs, *D*, containing a concentrated solution of potassium hydroxide, and a tube, *E*, containing soda lime in one leg and calcium chloride in the other. The

current of oxygen is passed through the tube, and the sulphur is very gently heated. The sulphur burns, forming sulphur dioxide, which is absorbed by the potash bulbs. Take care that the sulphur is all burnt, and that none is left sublimed in the cooler parts of the combustion tube. When all the sulphur has been oxidized, the apparatus is disconnected and reweighed. The increment in the weight of the potash bulbs represents the sulphur dioxide

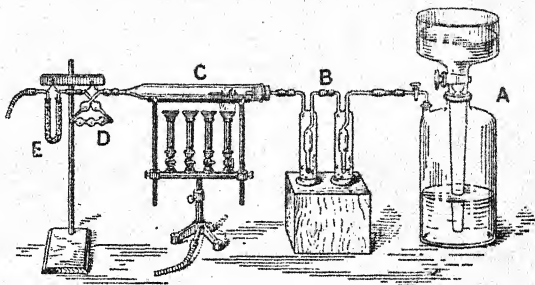


FIG. 168.—Synthesis of Sulphur Dioxide (by weight).

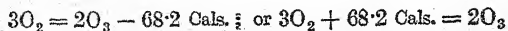
formed; and the loss in the weight of the porcelain boat, the amount of sulphur consumed. This proves that each of the different forms of sulphur is but a modification of one element.

A similar phenomenon is presented by many other elements—carbon, phosphorus, selenium, etc. That property in virtue of which one element may, without changing its state, exist in two or more forms with different properties is called *allotropy*—from the Greek *ἄλλος* (allos), another; *τρόπος* (tropos), shape. One allotropic form is an *alias*, so to speak, of the other. The less common form is sometimes called an “allotrope” or an “allotropic modification” of the other. The capacity of a substance to exist in two or more forms with different properties without changing the ultimate proportions of the constituents is not confined to elements; chemical compounds exhibit the same phenomenon, which may be called *isomerism* or *allotropism*. Are the different allotropic forms of an element to be regarded as one and the same substance? No. Each allotropic form has specific properties which distinguish it from all other substances. What, then, are allotropic states? asks W. S. Jevons, and answers: “Curious states which chemists frequently dispose of by calling them ‘allotropic’ a term frequently used when they are puzzled to know what has happened.”

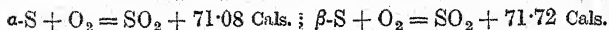
When a determination can be made of the molecular weight of two allotropic modifications, there is very frequently a difference, for the one modification contains more atoms per molecule than the other. This is the case, for instance, with oxygen and ozone. In consequence, it is often stated that allotropy is then due to a difference in the “molecular weight” of the element. This form of allotropy can be called *polytropism*. In other cases, it is assumed that the molecular weights are the same, as is probably the case with some of the different forms of sulphur, but the atoms of the molecule are arranged differently. This form of allotropy can be called *metatropism*. The idea is sometimes expressed in this way: “Just as bricks of the same kind in the hands of a builder may be fashioned into various structures, so Nature, from the same kind of atoms, builds up molecular structures with widely different properties.” The two graphic

formulae for the S_8 sulphur molecule indicated above have been suggested to account for the difference in the two crystalline forms of sulphur. The first one has been given to represent the structure of the molecule of β -sulphur; and the second the structure of α -sulphur. This, however, is pure hypothesis.

There is habitually a change of internal energy when one allotropic modification is changed into another. Ozone, it will be remembered, is formed from ordinary oxygen by an endothermal reaction. The heat absorbed in the production of one molecule of ozone is 34.1 Cal. (S. Jahn, 1908). Hence we write:



Accordingly, ozone is supposed to have more available energy than ordinary oxygen. Similar remarks might be applied to the different forms of carbon and of phosphorus. In the case of sulphur,



This means that the conversion of 32 parts by weight of rhombic sulphur into the monoclinic form is attended by an absorption of 0.64 Cal. There is a difference in the energy content of the two forms of sulphur as was the case with oxygen. Hence the definition: two allotropic modifications of a substance are composed of one element associated with different proportions of available or potential energy, and consequently they exhibit different physical and chemical properties. A definition of this kind describes the facts and no more. That is, of course, the function of a good definition. But there is a plausible finality about it not altogether pleasing. Energetics may tell us when reactions can occur, but it does not tell us how. Energy definitions in general are strictly non-committal and less likely to stimulate the imagination than views such as that which has just been styled "pure hypothesis." This indicates one great objection to the energetic method of dealing with chemical reactions. The atomic, molecular, and kinetic methods are far more likely to prompt new and fruitful investigations. However, if we accept J. C. Maxwell's view that the potential energy of a molecule depends essentially on the relative positions of the parts of the system in which it exists; and that potential energy cannot be transformed in any way without a change in the relative position of those parts, we see that the energy theory of allotropism is quite in harmony with the position theory, and that the one supplements the other. Changes of energy are dependent on changes in the configuration of a system. Chemists therefore seek answers to such questions as: What are the differences in the available energy of these molecules? What are the arrangements of the atoms in the different molecules? etc.

The transition of ordinary α -S to β -S is reversible. There is a definite transition temperature below or above which only one of the forms is stable, and the other form unstable. This is a case of enantiotropic allotropy—from the Greek *enantios* (enantios), opposite; *τροπος* (tropos), habit. The transformation of λ -S to μ -S is also reversible, but there is not a definite transition temperature, for the amount of each form present when the system is in equilibrium is determined by the temperature. The phenomenon is called dynamic allotropy, to distinguish it from that which precedes. In yet a third type of allotropy, the change is irreversible, one form is in a metastable condition at all temperatures. This is called monotropic allotropy, to distinguish it from the two phenomena which

precede. Examples—diamond and graphite, explosive antimony—will be considered later.

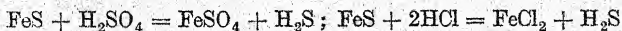
§ 8. Hydrogen Sulphide—Preparation and Occurrence.

Molecular weight, $\text{H}_2\text{S} = 34.09$. Melting point, -86° ; boiling point, -62° ; critical temperature, $+100^\circ$. Relative vapour density, 34.204 ($\text{H}_2 = 2$); 1.1895 (air = 1). 1000 c.c. weigh 1.5392 grams under standard conditions.

Occurrence.—Hydrogen sulphide occurs in several mineral waters (p. 178); in the exhalations from volcanic vents, etc. It is also formed during the putrefaction of animal and vegetable matters containing sulphur, and by the action of an aqueous solution of carbonic acid on sulphides which in turn are often formed by the reduction of sulphates by organic matter. Several bacteria during the putrefaction of organic matter produce hydrogen sulphide—the sulphur being derived from protean matters and from the reduction of sulphates—stimulated possibly by sunlight. This is said to occur at the mouths of rivers on the West Coast of Africa—*e.g.* the Congo river, which brings organic matter into intimate contact with sulphides. Hence, the atmosphere may be contaminated by hydrogen sulphide derived from these three courses. Hydrogen sulphide is readily oxidized and it cannot therefore persist very long in the atmosphere.

Historical.—Several references to hydrogen sulphide appear in the writings of the alchemists of the Middle Ages, where it is described under the general term “sulphurous vapour,” and their directions for the preparation of this product show that they referred to some fetid solutions of the polysulphides, which probably contained this gas, and were called “divine water”—from the Greek *θειον* (theion), divine or sulphurous. This liquid excited the attention of the alchemists, for it gave all kinds of colorations with solutions of the different metals; with vinegar, it gave a precipitate of sulphur and at the same time gave off a gas which blackened silver and some other metals. K. W. Scheele (1777) was the first to investigate the compound systematically.

Preparation.—Hydrogen and sulphur combine directly when sulphur vapour and hydrogen (or certain hydrocarbons) are passed through a red-hot tube, particularly if the tube be packed with pumice stone or some other similar porous material. Reintsch made the gas by melting suet and sulphur in a flask. The gas is best prepared by the action of dilute hydrochloric or sulphuric acid upon ferrous sulphide, FeS , which, in turn, is made by fusing iron and sulphur together (p. 32). The reactions are symbolized:



Hydrochloric acid is generally preferred to sulphuric acid because the resulting ferrous chloride— FeCl_2 —is not so liable to crystallize as ferrous sulphate— FeSO_4 . For small quantities, a similar apparatus to that employed for the preparation of hydrogen is used (Fig. 10); and for larger quantities, Kipp's apparatus may be used (Fig. 13). Comparatively large quantities of the gas are required intermittently in a testing laboratory, and scores of different forms of apparatus have been invented for the purpose. The gas is generally washed by passing it through a wash-bottle containing water.

Iron pyrites, FeS_2 , is scarcely attacked by hydrogen at high tempera-

tures, or by acids—hot or cold. If, however, the mineral be calcined to 500° – 800° , with about 15–20 per cent. of an alkali or alkaline earth, the product is then readily attacked and can be used for the preparation of the gas. Artificial ferrous sulphide, prepared as indicated above, generally contains a little free iron, and hence some hydrogen will be mixed with the gas. For ordinary purposes, this does not matter. The gas may also contain traces of hydrocarbons, etc., derived from the action of the acids on the impurities in the iron. A more pure gas is made by heating antimony sulphide with concentrated hydrochloric acid, and washing the gas in water; but calcium, barium, and magnesium sulphides also furnish as pure a gas at the ordinary temperatures: $\text{BaS} + 2\text{HCl} = \text{BaCl}_2 + \text{H}_2\text{S}$. Aluminium sulphide (*q.v.*) with water gives hydrogen sulphide.

Drying hydrogen sulphide.—If the gas is to be dried, dehydrated alumina or phosphoric oxide is used. Sulphuric acid should not be used because it is reduced by hydrogen sulphide and free sulphur is deposited: $\text{H}_2\text{S} + \text{H}_2\text{SO}_4 = \text{SO}_2 + 2\text{H}_2\text{O} + \text{S}$. Calcium chloride is often used for drying the gas in spite of the fact that there is a slight decomposition of the calcium chloride: $\text{CaCl}_2 + \text{H}_2\text{S} = \text{CaS} + 2\text{HCl}$, or $\text{CaCl}_2 + 2\text{H}_2\text{S} = \text{Ca}(\text{HS})_2 + 2\text{HCl}$.

§ 9. Hydrogen Sulphide—Properties and Composition.

The termination *uret*—from the Latin “it will burn”—was once applied to a series of inflammable compounds—*e.g.* sulphuretted hydrogen, phosphuretted hydrogen, carburetted hydrogen, etc. These terms are still used colloquially. Sulphuretted hydrogen, or hydrogen sulphide, is a colourless gas which smells like “ripe” eggs. The gas is very poisonous, and it produces headache and vomiting if breathed diluted with air, for a long time. According to L. J. Thénard, respiration in an atmosphere containing $\frac{1}{500}$ of its volume of hydrogen sulphide proved fatal to a dog. Inhalation of dilute chlorine obtained by wetting chloride of lime with acetic acid is recommended as an antidote. The gas has been used on certain metallurgical operations on a large scale—*e.g.* the purification of cadmium, and the separation of copper and nickel. Electrolytic processes, however, have largely displaced the use of this gas for operations of this character.

Action of cold.—The gas was liquefied by M. Faraday in 1823 by allowing pure ferrous sulphide and pure hydrochloric acid to act upon one another in a stout bent sealed glass tube. The gas condenses to a colourless limpid fluid at $+10^{\circ}$ under a pressure of 15 atmospheres. The liquid boils between -61° and -62° ; and it freezes at -85° . Liquid hydrogen sulphide forms a crystalline hydrate when heated with water in a sealed tube. The compound, probably $\text{H}_2\text{S} \cdot 6\text{H}_2\text{O}$ (R. de Forcrand, 1902), decomposes slowly at ordinary temperatures and pressures, but it may be preserved indefinitely in a sealed tube.

Aqueous solutions.—Hydrogen sulphide is fairly soluble in water; 100 volumes of water at 0° dissolve 437 volumes of the gas; and at 20° , 291 volumes. The solution is called “hydrogen sulphide water.” The solution is supposed to contain H^+ , and HS^+ , and S^+ ions, but the ionization, $\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^+$, is relatively much greater than $\text{H}_2\text{S} \rightleftharpoons 2\text{H}^+ + \text{S}^+$; and in a $\frac{1}{10}$ N-aqueous solution only 0.07 per cent. of the dissolved hydrogen

sulphide is supposed to be ionized. The gas can be expelled from the water by boiling. Owing to its solubility, the gas should not be collected over cold water, but it is sometimes collected over hot water. The gas is four times as soluble in alcohol as it is in water. The aqueous or alcoholic solution decomposes slowly with deposition of sulphur, particularly if exposed to the light— $2\text{H}_2\text{S} + \text{O}_2 = 2\text{H}_2\text{O} + \text{S}$. The aqueous solution has an acid reaction, and it reddens blue litmus. Hydrogen sulphide reacts with bases, forming sulphides. When it is desired to emphasize the acid nature of the gas, the aqueous solution is sometimes called hydro-sulphuric acid.

Hydrogen sulphide is inflammable in air, and burns with a bluish flame, forming sulphur dioxide and water: $2\text{H}_2\text{S} + 3\text{O}_2 = 2\text{SO}_2 + 2\text{H}_2\text{O}$. A mixture of two volumes of hydrogen sulphide with three volumes of oxygen explodes violently when ignited. If the supply of air is limited, free sulphur may be formed: $2\text{H}_2\text{S} + \text{O}_2 = 2\text{H}_2\text{O} + 2\text{S}$. A lighted taper dipped into a jar of the gas shows that the gas does not support combustion.

Decomposition of hydrogen sulphide.—The thermal value of the reaction between hydrogen and sulphur is small: $\text{H}_2 + \text{S}_{\text{gas}} = \text{H}_2\text{S} + 4.82 \text{ Cals.}$ The gas is

easily dissociated by passing it through a hot porcelain tube. The dissociation begins about 400° , and it is complete at about 1700° . It will be remembered that hydrogen sulphide was formed by passing hydrogen and sulphur vapour through a porcelain tube. This means that the reaction belongs to the type of opposing reactions—discussed on p. 119. This reaction is

symbolized: $2\text{H}_2\text{S} \rightleftharpoons 2\text{H}_2 + \text{S}_2$. Hydrogen sulphide is also decomposed by passing electric sparks through the gas confined in a tube, Fig. 169, over mercury. No change in volume occurs, but the decomposition of the gas is evidenced by the deposition of sulphur on the glass in the vicinity of the sparks. There is a considerable local rise of temperature in the vicinity of the sparks, but the gas is immediately cooled as it diffuses into the surrounding gas. The products of the reaction do not then have time to react in the converse direction. The effect can be compared with a similar result obtained with Deville's "hot and cold tube" (p. 229). Decomposition is complete because the sulphur is removed from the zone of the reaction, and deposited as a solid about the anode, the hydrogen about the cathode. When hydrogen sulphide is passed through a hot tube, decomposition is not complete, except at very high temperatures, because the sulphur is present in the reacting system as a vapour, and it sets up a back reaction with the hydrogen, reforming hydrogen sulphide.

Reducing action.—The relatively small amount of energy absorbed when hydrogen sulphide decomposes corresponds with the fact that it is

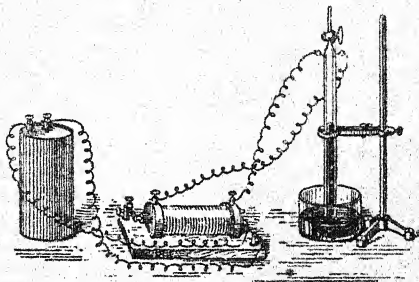


FIG. 169.—The Decomposition of Hydrogen Sulphide by Electric Sparks.

readily decomposed and the products of decomposition exert a powerful reducing action. The action of the gas on sulphuric acid has just been indicated. Hydrogen sulphide reduces moist sulphur dioxide with the separation of sulphur: $2\text{H}_2\text{S} + \text{SO}_2 = 2\text{H}_2\text{O} + 3\text{S}$. Oxygen, and substances which easily part with their oxygen, oxidize the hydrogen and liberate sulphur, or form sulphur dioxide. The turbidity observed in aqueous solutions of the gas which have stood some time is due to the separation of sulphur and the oxidization of the hydrogen by air. Iodic, bromic, chloric, and chromic acids are readily reduced by hydrogen sulphide. Fuming nitric acid is reduced with explosive violence. This can be shown by dropping acid into a jar of the gas. Chlorine, bromine, and fluorine also decompose hydrogen sulphide with the separation of sulphur, and the formation of the respective halogen compounds. This can be proved by bringing a jar of chlorine and a jar of hydrogen sulphide mouth to mouth. Hydrochloric acid will be formed: $\text{H}_2\text{S} + \text{Cl}_2 = \text{S} + 2\text{HCl}$. This property is utilized in making hydriodic and hydrobromic acids. A piece of bright silver is very quickly blackened when exposed to the gas owing to the formation of silver sulphide. Hydrogen sulphide is sometimes present in the atmosphere; hence silver often tarnishes when exposed to the air of towns. The tarnishing of silver by hydrogen sulphide is illustrated by the use of silver spoons with eggs. Tin and lead are also quickly tarnished by the gas. Some metals decompose hydrogen sulphide very quickly under the influence of heat. Tin, lead, and cadmium are examples.

Sulphides of the metals are formed, and hydrogen gas is liberated: $\text{Sn} + \text{H}_2\text{S} = \text{SnS} + \text{H}_2$.

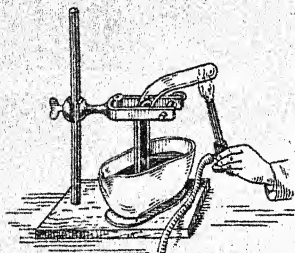


FIG. 170.—Composition of Hydrogen Sulphide.

Composition and formula.—If a known volume of hydrogen sulphide be heated with metallic tin in a tube over mercury—Fig. 170¹—tin sulphide and free hydrogen equal to the original volume of hydrogen sulphide are formed; similarly, hydrogen sulphide, when decomposed by electric sparks—Fig. 169—suffers no change in volume. Hence, from Avogadro's hypothesis, it follows at once that one molecule of hydrogen sulphide

contains one molecule—two atoms—of hydrogen, and that the formula of hydrogen sulphide is H_2S_n , where n has not been determined. The specific gravity of hydrogen sulphide (air = 1) is 1.1912; that is, if $\text{H}_2 = 2$ be the unit, the relative density of hydrogen sulphide is $1.1912 \times 28.755 = 34.253$. Hence—

Molecular weight of hydrogen sulphide	34.204
Weight of hydrogen in the molecule	2.016
Weight of sulphur in the molecule	32.188

This result is sufficiently close to the atomic weight of sulphur — 32.07, to prove that there can be one and only one atom of sulphur in the molecule. It is therefore concluded that the formula of hydrogen sulphide is H_2S .

¹ A slight depression in the tube retains the tin when the mercury is being displaced by the gas while charging the tube.

§ 10. Sulphides.

The sulphides can be regarded as salts of hydrosulphuric acid even though they are not always prepared directly from hydrogen sulphide. Many sulphides are made by the direct union of sulphur with the metals (p. 32). Hydrogen sulphide also forms either sulphides or hydrosulphides with the oxides and hydroxides of the metals, for instance, with lead oxide: $\text{PbO} + \text{H}_2\text{S} = \text{PbS} + \text{H}_2\text{O}$; and with lead salts: $\text{PbCl}_2 + \text{H}_2\text{S} = \text{PbS} + 2\text{HCl}$. The lead sulphide is dark brown. Hence paints containing lead compounds, when exposed to air contaminated with hydrogen sulphide, are "blackened." If thoroughly dried, hydrogen sulphide has no appreciable action on most of the oxides—lead, iron, antimony, magnesium, barium, etc.—at ordinary temperatures. Many of the metallic sulphides prepared by the action of hydrogen sulphide on metal salt solutions have characteristic colours; and, in consequence, the colour of the sulphide precipitated when hydrogen sulphide is passed into a solution of salt of the metal, is often strong circumstantial evidence of the presence of particular metals. Hydrosulphides are formed by the action of hydrogen sulphide on some of the hydroxides: $\text{KOH} + \text{H}_2\text{S} = \text{KSH} + \text{H}_2\text{O}$. The group "SH" resembles the group "OH" in that it is a monad radicle, and forms a group of compounds called hydrosulphides.

Polysulphides.—Sulphur is fairly soluble in aqueous solutions of the soluble sulphides, forming a group of polysulphides. With sodium, for example, a series of sulphides ranging from sodium monosulphide, Na_2S , to the pentasulphide, Na_2S_5 , can be obtained. The composition and the relations of the polysulphides have not been clearly demonstrated. Aqueous solutions of ammonia saturated with hydrogen sulphide mainly consist of ammonium hydrosulphide, NH_4HS (formed: $\text{NH}_4\text{OH} + \text{H}_2\text{S} = \text{NH}_4\text{HS} + \text{H}_2\text{O}$), some ammonium sulphide, $(\text{NH}_4)_2\text{S}$, and ammonium hydroxide. The mixture is often called "ammonium sulphide." When "ammonium sulphide" is allowed to stand, the solution is oxidized, and free sulphur is formed. This dissolves in the "ammonium sulphide," forming a polysulphide, $(\text{NH}_4)_2\text{S}_n$, where n may be anything from 1 to 9. The yellow solution—called "yellow ammonium sulphide"—is a reagent much used in analytical work.

The alkali sulphides.—The sulphides are analogous in many respects with the oxides, and, as with the oxides, we have basic, acidic, and neutral sulphides, as well as persulphides. Sulphur is less acidic (electro-negative) than oxygen and chlorine, and accordingly, the compounds of sulphur and oxygen and chlorine are not sulphides of oxygen and chlorine, but oxides and chlorides of sulphur. The anhydrous sulphides are made by the action of sulphur vapour on an excess of the alkali metal between 200° and 300° in vacuo. The excess of metal can be distilled off, and the white crystalline sulphide remains. Just as the metal sodium decomposes water, H_2O , forming sodium hydroxide, NaOH , so does the metal sodium when heated with hydrogen sulphide, H_2S , form the sulphur analogue of sodium hydroxide, namely, sodium hydrosulphide, NaSH . The same substance is formed when alkaline hydroxide is saturated with a solution of hydrogen sulphide, and, in the case of potassium hydroxide, a crystalline monohydrate, $2\text{KSH} \cdot \text{H}_2\text{O}$, can be isolated from the solution. The aqueous

solutions are hydrolyzed and hence anhydrous products cannot be obtained in such solutions excepting perhaps the rubidium and caesium compounds. When sodium ethoxide—formed by the action of sodium on ethyl alcohol—is saturated with hydrogen sulphide, and then treated with ether, fairly pure sodium hydrosulphide is formed. When potassium hydroxide and potassium hydrosulphide are mixed in equivalent proportions, potassium monosulphide and water are formed: $\text{KOH} + \text{KSH} = \text{K}_2\text{S} + \text{H}_2\text{O}$. Reddish-white prismatic crystals of the pentahydrate, $\text{K}_2\text{S} \cdot 5\text{H}_2\text{O}$, can be obtained by the evaporation of the aqueous solution in vacuo. In the case of sodium, $\text{Na}_2\text{S} \cdot 9\text{H}_2\text{O}$ is formed. The alkaline sulphides are at once hydrolyzed by water: $\text{K}_2\text{S} + \text{H}_2\text{O} \rightleftharpoons \text{KSH} + \text{KOH}$. The solution quickly oxidizes on exposure to the air; sulphur is first liberated, and this is dissolved by an alkali sulphide, forming a polysulphide. The polysulphide on further oxidation forms a colourless solution of the thiosulphate. Ozone reacts with sodium sulphides, forming sodium sulphate and a little thiosulphate; with sodium hydrosulphide, free hydrogen sulphide is formed as well. When potassium carbonate or hydroxide and sulphur are heated together a mixture of variable composition containing sulphate, thiosulphate, and polysulphide is formed—probably $6\text{KOH} + 8\text{S} = 2\text{K}_2\text{S}_3 + \text{K}_2\text{S}_2\text{O}_3 + 3\text{H}_2\text{O}$ represents the primary reaction, and the potassium trisulphide, K_2S_3 , dissolves sulphur, forming potassium pentasulphide, K_2S_5 , with potassium tetrasulphide, K_2S_4 , as an intermediate product. The reddish-brown solution of sulphur in potassium hydroxide or carbonate is called “liver of sulphur,” or “hepar sulphuris.” Some sulphate and sulphite are also formed as oxidation products, and the solution accordingly contains these salts as well as the polysulphides and thiosulphate. Sulphur readily reacts with alkali by hydrosulphide in alcoholic solution, forming in the case of sodium, sodium tetrasulphide, Na_2S_4 , and sodium disulphide, Na_2S_2 , when the right proportions of sulphur are employed.

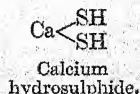
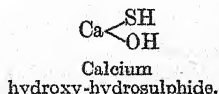
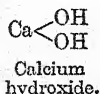
Copper sulphides.—Cuprous sulphide, Cu_2S , is produced when copper burns in sulphur vapour; when an excess of copper filings is heated with sulphur; and, as a black precipitate, when hydrogen sulphide is passed through solutions of cuprous salts: $2\text{CuCl} + \text{H}_2\text{S} = \text{Cu}_2\text{S} + 2\text{HCl}$. The cuprous sulphide is soluble in warm dilute nitric acid, forming copper nitrate and sulphur. Cuprous sulphide occurs native in rhombic crystals as *copper glance*. A bluish mass of cupric sulphide is formed when copper or cuprous sulphide is heated with sulphur to a temperature below 114° ; and as a black precipitate when hydrogen sulphide is passed into solutions of cupric salts. The precipitate is inclined to run through filter paper because it is a hydrosol (colloid). The precipitate is coagulated by the addition of dilute hydrochloric acid. Cupric sulphide is soluble in hot dilute nitric acid; insoluble in dilute sulphuric acid (cadmium sulphide is soluble under the same conditions); insoluble in potassium or sodium sulphide, and insoluble in ammonium sulphide. It also dissolves in potassium cyanide, forming a complex cyanide from which hydrogen sulphide does not precipitate copper.

Silver sulphide, Ag_2S .—Silver sulphide is precipitated by hydrogen sulphide from neutral, acid, or ammoniacal solutions. The precipitate is “insoluble” in ammonia, alkaline sulphides, and in dilute potassium cyanide; but it is soluble in concentrated potassium cyanide and in hot dilute nitric acid. Silver sulphide melts at about 812° . Fine crystals

can be prepared by heating the precipitate with ammonium sulphide in a sealed tube between 150° and 200° . If the silver nitrate solution be highly concentrated, a yellowish green precipitate is formed which, when washed with dilute nitric acid and dried, has the empirical composition $\text{Ag}_2\text{S} \cdot \text{AgNO}_3$, silver sulphonitrate. Water decomposes the compound into soluble silver nitrate and insoluble silver sulphide.

Gold sulphide.—Hydrogen sulphide precipitates in the cold a black gold disulphide, Au_2S_2 : The precipitate is insoluble in dilute acids; readily soluble in *aqua regia*—that is, a mixture of nitric and hydrochloric acids—forming auric chloride. Gold disulphide is also soluble in ammonium sulphide, but more readily in potassium sulphide, forming potassium thioaurite, $\text{S}=\text{Au}-\text{SK}$, from which it is said yellowish-brown gold trisulphide, Au_2S_3 , can be precipitated by hydrogen sulphide. Metallic gold is precipitated from hot solutions of auric chloride by hydrogen sulphide.

Alkaline earth sulphides.—The monosulphides are formed by reducing the sulphates with carbon. In the case of barium, this reaction is the starting-point for the manufacture of the barium salts from barytes, BaSO_4 . The reaction is symbolized: $\text{BaSO}_4 + 4\text{C} = \text{BaS} + 4\text{CO}$. Calcium sulphide, CaS , is formed by the action of hydrogen sulphide on heated lime: $\text{Ca}(\text{OH})_2 + \text{H}_2\text{S} = 2\text{H}_2\text{O} + \text{CaS}$. The alkaline earth sulphides are white or yellowish-white, and phosphoresce in the dark, after exposure to daylight. These sulphides are almost insoluble in water, but like the alkali sulphides, they are hydrolyzed very quickly in boiling water, probably forming hydrosulphide and hydroxide:



The hydrosulphides may be obtained in aqueous solution by the action of an excess of hydrogen sulphide on the hydroxide, as indicated above, and crystals of $\text{Ca}(\text{SH})_2 \cdot 6\text{H}_2\text{O}$ can be obtained from the solution. These crystals, when heated in a stream of hydrogen sulphide, give calcium monosulphide, CaS . Just as hot alkaline hydrosulphides dissolve sulphur, forming alkali polysulphides, so does boiling milk-of-lime and sulphur form calcium polysulphides. According to H. V. Tartar (1914), calcium tetrasulphide is one of the first products. The primary reaction is represented: $3\text{Ca}(\text{OH})_2 + 10\text{S} = 2\text{CaS}_4 + \text{CaS}_2\text{O}_3 + 3\text{H}_2\text{O}$. The thio-sulphate, CaS_2O_3 , decomposes into sulphur and sulphite. It reacts with an excess of sulphur, forming calcium pentasulphide, CaS_5 . The sulphides of the alkaline earths—calcium, barium, strontium, and magnesium—can be used as a source of pure hydrogen sulphide, since they decompose with acids giving hydrogen sulphide: $\text{CaS} + 2\text{HCl} = \text{H}_2\text{S} + \text{CaCl}_2$. Calcium sulphide is a by-product in Leblanc's process (*q.v.*). *Canton's phosphorus* was made by Canton (1768) by heating oyster shells with sulphur.

The commercial sulphides of the alkaline earths after exposure to light appear luminous when placed in the dark. The feeble light emitted by these substances gradually diminishes in intensity, but the property is recovered on exposure to light. Calcium sulphide, CaS , was formerly termed "*Canton's phosphorus*," and barium sulphide, BaS , "*Bononian*

(i.e. Bolognian) *phosphorus*." What is here called barium sulphide was discovered by V. Cascariolo, of Bologna, towards the end of the 16th century, while he was experimenting on the action of heat on a mixture of barytes and charcoal. He thought the product of the action "was capable of attracting and retaining the golden light of the sun," and for a time the substance was called *lapis solaris* (sun-stone). These substances are now used in the manufacture of "*luminous paints*." The pure sulphides do not phosphoresce, and the property is therefore dependent on the presence of some foreign substance. Minute traces of other elements, bismuth, cadmium, manganese, zinc, etc., modify the colour of the phosphorescent glow.

Zinc sulphide.—Zinc sulphide is formed as a white amorphous precipitate when an alkaline sulphide is added to a solution of zinc salt, or when hydrogen sulphide is passed through an alkaline solution of a zinc salt. Zinc sulphide melts at about 1660° . It is not dissolved by organic acids like acetic and formic acids, but it is soluble in mineral acids evolving hydrogen sulphide. Hence zinc sulphide is not precipitated by hydrogen sulphide in acid solutions; cadmium sulphide is precipitated in acid solutions. This subject is discussed in the next section. There are two native zinc sulphides, the one, *sphalerite* or ordinary *zinc blende*, is stable below 1020° , and the other, *wurtzite*, is stable above this temperature. *Sphalerite* has a specific gravity 4.090 and crystallizes in the cubic system; *wurtzite* has a specific gravity 4.087, and crystallizes in the hexagonal system. Cadmium sulphide varies in tint from a bright yellow to an orange-red, according to the temperature of precipitation, nature of solution, etc. If hydrogen sulphide be passed through a solution of cadmium chloride, the precipitate which forms is an intense orange-red colour owing to the formation of cadmium thiochloride, Cd_2SCl_2 . The reaction is represented: $2\text{CdCl}_2 + \text{H}_2\text{S} = 2\text{HCl} + \text{Cl}-\text{Cd}-\text{S}-\text{Cd}-\text{Cl}$. The thiochloride passes into the sulphide CdS by the continued action of the hydrogen sulphide, and at the same time, the tint of the precipitate becomes lighter in colour. With the chloride, the precipitate never attains the sulphur-yellow tint formed when solutions of cadmium nitrate or sulphate are employed. *Greenockite* is the native form of cadmium sulphide, and occurs in hexagonal crystals. Cadmium sulphide is used as a pigment for oil and water colours. Cadmium sulphide is insoluble in ammonium sulphide; arsenic, which also forms a bright yellow precipitate, is soluble in ammonium sulphide. The fact that zinc oxide dissolves in alkalis while zinc sulphide does not, illustrates the stronger acidic qualities of oxygen in contrast with sulphur.

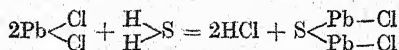
Mercury sulphides.—It is said that mercurous sulphide, Hg_2S , is produced in the form of brownish-black plates by the prolonged action of cold concentrated sulphuric acid on mercury. There appears to be some doubt about the existence of mercurous sulphide. When hydrogen sulphide acts upon mercurous salts, a mixture of mercuric sulphide and mercury results. Mercuric sulphide, HgS , is made by rubbing mercury and sulphur together in a mortar. It is also formed as a black precipitate by the action of hydrogen sulphide upon a mercuric salt. Mercuric sulphide is insoluble in ammonium sulphide and in alkaline hydroxides, but it dissolves in concentrated solutions of alkaline sulphides, more particularly the polysulphides, forming thio-salts: $\text{HgS} + \text{K}_2\text{S} = \text{Hg}(\text{SK})_2$.

The thio-salt is completely hydrolyzed by water, forming red mercuric sulphide: $\text{Hg}(\text{SK})_2 + \text{H}_2\text{O} = \text{KOH} + \text{KSH} + \text{HgS}$. When hydrogen sulphide is first passed through the solution of the mercuric salt, a white precipitate is formed. This is supposed to be mercury thiochloride, $\text{Cl}-\text{Hg}-\text{S}-\text{Hg}-\text{S}-\text{Hg}-\text{Cl}$. The mercury thiochloride gradually turns brown and then black if the current of gas is continued: $\text{Hg}_2\text{S}_2\text{Cl}_2 + \text{H}_2\text{S} = 3\text{HgS} + 2\text{HCl}$. The black precipitate is almost insoluble in boiling dilute acids though hot concentrated nitric acid gradually converts it into a white mercury thionitrate, $\text{Hg}_2\text{S}_2(\text{NO}_3)_2$, and finally into soluble mercuric nitrate. There are three different forms of crystalline mercuric sulphide; two of these occur in nature. The one is black *metacinnabar* with a specific gravity 7.7 and crystallizes in the cubic system; it can be prepared artificially by treating a dilute acid solution of mercuric salts with sodium thiosulphate; by treating a concentrated solution in the same manner, a scarlet mercuric sulphide is formed with a specific gravity 7.2. This does not occur in nature. Annabar is obtained from alkaline solutions, metacinnabar from acidic solutions. *Cinnabar*, *vermilion*, or native sulphide of mercury, HgS , is red. If the black sulphide be sublimed, a red crystalline sulphide is formed. The red crystalline variety of mercuric sulphide is more stable than the black variety. When either the black or the red variety is heated to about 400° , the black compound is formed on cooling, and this may be transformed into the red variety by merely scratching the surface. Zinc, cadmium, and mercuric sulphides can be sublimed without melting; the sublimation begins respectively at 1180° , 980° , and 446° . Vermilion is used as a pigment in oil-painting; it is made artificially by the action of ammonium or alkali sulphides on black mercuric sulphide.

Tin sulphides.—When tinfoil is burned in sulphur vapour stannous sulphide, SnS , is formed. Stannous hydrosulphide, $\text{Sn}(\text{HS})_2$, is precipitated as a brown powder when hydrogen sulphide is passed through a solution of the stannous salt: $\text{SnCl}_2 + 2\text{H}_2\text{S} = 2\text{HCl} + \text{Sn}(\text{SH})_2$. The brown precipitate becomes black and anhydrous on drying. The precipitate dissolves in concentrated hydrochloric acid and consequently no tin is precipitated if hydrogen sulphide be passed through a strongly acid solution; on diluting such a solution, however, the stannous sulphide is precipitated. Unlike arsenic sulphide, tin sulphide is insoluble in ammonia and ammonium carbonate, and in colourless ammonium sulphide; but it is readily soluble in ammonium and alkaline polysulphides, forming thioannates, e.g. potassium thioannate, K_2SnS_3 , thus: $\text{SnS} + \text{K}_2\text{S} + \text{S} \rightarrow \text{SSn}(\text{SK})_2$. If the solution be acidified, yellow stannic sulphide is precipitated: $\text{S}=\text{Sn}=(\text{SK})_2 + 2\text{HCl} = 2\text{KCl} + \text{H}_2\text{S} + \text{SnS}_2$. Stannic sulphide, SnS_2 , is precipitated by passing hydrogen sulphide through a (not too acid) solution of a stannic salt, e.g. $\text{SnCl}_4 + 2\text{H}_2\text{S} = 4\text{HCl} + \text{SnS}_2$. The sulphide is soluble in hydrochloric acid, and hence no precipitation of the sulphide occurs if the solution be strongly acid; such a solution, saturated with hydrogen sulphide, precipitates stannic sulphide when diluted. The yellow stannic sulphide appears to be the anhydride of a thio-stannic acid, H_2SnS_3 , for it dissolves in alkali sulphides, forming soluble thioannates as indicated above. Stannic sulphide is insoluble in ammonia and ammonium carbonate; and it is converted into the oxide by roasting in air. Stannous sulphide alone is prepared in the dry way by heating

tin and sulphur together because the heat developed during the reaction converts the stannic sulphide into stannous sulphide and sulphur. Stannic sulphide can be made in a dry way by heating tin amalgam, ammonium chloride and sulphur in a retort. A complex reaction takes place, resulting in the formation of a mass of yellow scales called "mosaic gold," and this is used as a pigment. "Mosaic gold" is not attacked by alkaline sulphides nor by nitric acid; it is attacked by *aqua regia*, forming stannic chloride and sulphur. The insoluble sulphides of tin are most readily obtained in a soluble condition by fusing them together with sodium carbonate and sulphur, and extracting the resulting sodium thioannate with water.

Lead sulphide, PbS.—Lead sulphide occurs in nature as galena in well-formed cubic crystals with a lustre resembling metallic lead. Lead sulphide is formed by reducing the sulphate with carbon, by heating lead in sulphur vapour, and as a black precipitate by passing hydrogen sulphide through neutral, acid, or alkaline solutions of a lead salt. If hydrochloric acid be present, an orange, yellow, or red precipitate of lead thiocchloride, Pb_2SCl_2 , may be formed:



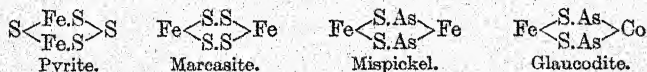
This is immediately decomposed by more hydrogen sulphide, forming a black lead sulphide; with solutions of lead bromide, lead sulphobromide, Pb_2SBr_2 , is formed. These compounds are the sulphur analogues of lead oxychloride, Pb_2OCl_2 . Boiling dilute nitric acid dissolves lead sulphide, forming lead nitrate with the separation of sulphur. Concentrated nitric acid oxidizes it to lead sulphate. Unlike tin sulphide, lead sulphide is insoluble in alkaline hydroxides and sulphides. Synthetic lead sulphide melts at 1112° , and galena at a lower temperature; lead sulphide begins to sublime at about 950° in vacuo or in a current of an inert gas, forming small cubic crystals. Heated with free access of air, it forms lead sulphate.

Aluminium sulphide, Al_2S_3 .—Aluminium sulphide can be prepared as a greyish black solid by heating finely divided aluminium with iron pyrites: $4\text{Al} + 3\text{FeS}_2 = 2\text{Al}_2\text{S}_3 + 3\text{Fe}$; or by heating a mixture of alumina, carbon, and sulphur at a white heat: $\text{Al}_2\text{O}_3 + 3\text{C} + 3\text{S} = \text{Al}_2\text{S}_3 + 3\text{CO}$; it is also formed when sulphur is thrown upon strongly heated aluminium; or better, by heating the metal in sulphur vapour. Aluminium sulphide is decomposed by water with the evolution of hydrogen sulphide: $\text{Al}_2\text{S}_3 + 6\text{H}_2\text{O} = 2\text{Al}(\text{OH})_3 + 3\text{H}_2\text{S}$. Atmospheric moisture also decomposes the sulphide in a similar manner. The sulphides of aluminium and of chromium cannot be prepared in the presence of water; hence when ammonium sulphide is added to solutions of chromium or aluminium salts, the hydroxides, not the sulphides, are precipitated. The hydrosulphide is probably formed first: $\text{AlCl}_3 + 3\text{NH}_4\text{HS} = 3\text{NH}_4\text{Cl} + \text{Al}(\text{SH})_3$; and this is at once hydrolyzed: $\text{Al}(\text{SH})_3 + 3\text{H}_2\text{O} = \text{Al}(\text{OH})_3 + 3\text{H}_2\text{S}$.

Iron sulphides.—By projecting a mixture of iron filings and sulphur into a red-hot crucible, a fused mass of ferrous sulphide, FeS , is formed. A little black ferrous sulphide is precipitated when hydrogen sulphide is passed through neutral solutions of ferrous salts, if an alkaline acetate be present, a little more ferrous sulphide is precipitated, but the precipitation is not complete; alkali and ammonium sulphide precipitate black

ferrous sulphide: $\text{FeCl}_2 + (\text{NH}_4)_2\text{S} = 2\text{NH}_4\text{Cl} + \text{FeS}$. The precipitate is readily soluble in dilute acids, even acetic acid, with the evolution of hydrogen sulphide. Moist ferrous sulphide is readily oxidized when exposed to the air, forming first a brownish basic sulphate with the separation of sulphur. Hydrogen sulphide and ammonium sulphide readily reduce ferric to ferrous salts with the separation of sulphur: $2\text{FeCl}_3 + \text{H}_2\text{S} = 2\text{FeCl}_2 + 2\text{HCl} + \text{S}$; and the ferrous chloride then behaves as indicated above. Hence ferric sulphide, Fe_2S_3 , cannot be made by precipitation with hydrogen sulphide, but it can be made by fusing equal weights of iron and sulphur at about 550° ; and by passing a current of hydrogen sulphide over ferric oxide heated to about 100° .

Ferrous sulphide is comparatively rare in nature, while iron pyrites, FeS_2 , is exceedingly common. The latter—iron disulphide, FeS_2 —can be made artificially by gently heating iron with an excess of sulphur at a low red heat—about 500° ; at 700° it loses sulphur, forming ferrous sulphide, FeS . The term *pyrites* is derived from *πῦρ*, fire, in allusion to the production of sparks when a piece is sharply struck. Iron pyrites occurs in two forms; one, *pyrite*, has a specific gravity about 5.1. It crystallizes in the cubic system, is but very slowly oxidized in air; the other, *marcasite*, has a specific gravity about 4.8, crystallizes in the rhombic system, is less stable than pyrite, and is oxidized comparatively quickly in air, and on this account marcasite has been largely used in the manufacture of “copperas,” ferrous sulphate. The former has been prepared artificially; the latter has not. The stable sulphides—cinnabar, sphalerite, and pyrite—are obtained by crystallization from alkaline solutions, and the unstable forms—metacinnabarite, wurtzite, and marcasite—from acidic solutions. Some sulphur is made by the dry distillation of iron pyrites—the reaction is symbolized $3\text{FeS}_2 = \text{Fe}_3\text{S}_4 + \text{S}_2$; and attention is usually drawn to the analogous reaction $3\text{MnO}_2 = \text{Mn}_3\text{O}_4 + \text{O}_2$, although the formation of Fe_3S_4 has not been established with certainty. Pyrite is not acted upon by dilute acids, but hot concentrated hydrochloric acid decomposes it, forming hydrogen sulphide and sulphur. If heated in hydrogen, sulphur is evolved and ferrous sulphide remains. When pyrite is heated with hydrogen dioxide, one atom of sulphur is liberated for every three atoms oxidized to sulphuric acid, while with marcasite, half the sulphur is oxidized to sulphuric acid, and half is liberated as sulphur. Hence, E. Arbeiter (1913) represents the molecules:



When heated in air, pyrites and many other sulphides of iron produce ferric oxide and sulphur dioxide. This action is partly due to the stronger acidic properties of oxygen over sulphur, and also to the volatility of sulphur dioxide which removes sulphur from the zone of the reaction. Many sulphides, e.g. lead sulphides, form sulphates when heated in air. *Magnetic iron pyrites*, or ferrosferric sulphide, also called *pyrrhotine*, or *pyrrhotite*, supposed to be in the ideal case: Fe_7S_{11} , is the sulphur analogue of magnetic oxide of iron, Fe_3O_4 . The ratio Fe : S varies a great deal in native specimens; their composition is said to “range from Fe_7S_6 to $\text{Fe}_{18}\text{S}_{17}$, chiefly $\text{Fe}_{11}\text{S}_{12}$.” This is taken to mean that the minerals are

probably mixtures of different sulphides, and may be of sulphur, etc. There is some difference of opinion as to whether the crystals of pyrrhotite are monoclinic or rhombic. Since the sulphur sublimes in vacuo from pyrrhotite at 530° , while free sulphur sublimes in vacuo at 60° to 70° , and half the sulphur passes from pyrite at 430° , pyrrhotite cannot be either a mixture of pyrite and ferrous sulphide, or a solid solution of sulphur in ferrous sulphide, but is rather a solid solution of ferrous sulphide and an iron compound rich in sulphur. Nickel sulphide melts at 797° ; cobalt sulphide at 1100° , and manganese sulphide at 1620° .

§ 11. The Action of Hydrogen Sulphide on Metallic Salt Solutions.

Justus von Liebig raised hydrogen sulphide to the dignity of a group reagent in the analysis of complex mixtures. As a group reagent it plays so important a rôle that all attempts to displace it by another reagent have not been successful. Wherever chemical investigations are carried on it is well-nigh indispensable.—E. GOLDSCHMIDT.

Hydrogen sulphide is a valuable reagent. Its reactions with the different metallic salts enable the metals to be separated into groups as a preliminary to more detailed examination. Thus—

I. Sulphides insoluble in dilute acids.

(a) Soluble in alkaline sulphides—arsenic, antimony, stannic, gold, germanium, molybdenum, tellurium, tungsten, iridium, and platinum sulphides.

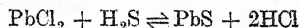
(b) Insoluble in alkaline sulphides—mercury, silver, lead, copper, bismuth, cadmium, and stannous sulphides.

II. Sulphides soluble in dilute mineral acids, but insoluble in the presence of alkalis—iron, cobalt, nickel, manganese, and zinc sulphides.

III. Sulphides not precipitated by hydrogen sulphide—chromium, aluminium, magnesium, barium, strontium, calcium, potassium, and sodium. Chromium and aluminium are precipitated as hydroxides.

The method of classifying certain elements into groups—those which form soluble and those which form insoluble sulphides in hydrochloric acid—frequently conveys wrong ideas of the properties of the sulphides. The solubility of the sulphides depends upon the concentration of the acid. For instance, if hydrogen sulphide be passed into 5 c.c. of a solution of 2 grams of tartar emetic—potassium antimonyl tartrate—in 15 c.c. of hydrochloric acid (sp. gr. 1.175) and 85 c.c. of water, antimony sulphide will be precipitated, but not if 15 c.c. of hydrochloric acid had been employed without the water. In one case, $2\text{SbCl}_3 + 3\text{H}_2\text{S} = \text{Sb}_2\text{S}_3 + 6\text{HCl}$; and in the second case, $\text{Sb}_2\text{S}_3 + 6\text{HCl} = 3\text{H}_2\text{S} + 2\text{SbCl}_3$. In other words, the antimony sulphide, in the second case, is decomposed by the acid as fast as it is formed. Similarly, no lead will be precipitated by hydrogen sulphide from a solution containing 3 per cent. of hydrochloric acid, HCl; and if the solution has 2.5 per cent. of acid, the lead sulphide will be imperfectly precipitated—i.e. part will be precipitated, and part will be decomposed as fast as it is formed. Similarly, a 5 per cent. boiling solution of hydrochloric acid will prevent the precipitation of cadmium sulphide.

If a metallic sulphide, MS, be treated with hydrochloric acid, hydrogen sulphide and a metallic chloride will be formed: $MS + 2HCl = MCl_2 + H_2S$. Conversely, when a metallic chloride in aqueous solution is treated with hydrogen sulphide, the metallic sulphide and hydrochloric acid will be produced: $MCl_2 + H_2S = MS + 2HCl$. Hydrochloric acid thus accumulates in the solution as the action goes on. If any more sulphide be produced, after the hydrochloric acid has attained a certain limiting concentration, the excess of sulphide will be at once decomposed by the acid. There are then two simultaneous opposing reactions: (1) Formation of the metallic sulphide and hydrochloric acid; and (2) formation of chloride and hydrogen sulphide. In illustration, if a current of hydrogen sulphide be passed through a saturated solution of zinc chloride, part of the metal is precipitated, but when the hydrochloric acid has attained a certain concentration, the action apparently ceases because the reverse change sets in. Hence the precipitation will be incomplete. In illustration, take the case of lead chloride:



When equilibrium is established, the solution contains lead chloride, hydrogen sulphide, and hydrogen chloride. Using symbols in square brackets to represent the concentrations (gram-molecules per litre) of the respective compounds in the solution, it follows from the equilibrium law, that:

$$\frac{[PbCl_2] \times [H_2S]}{[HCl]^2} = \text{Constant}$$

This shows that if the concentration of the acid be increased, and the concentration of the hydrogen sulphide be constant, the amount of lead chloride which remains in solutions (that is, escapes precipitation) will increase in order to keep the numerical value of the "constant" always the same. Conversely, if it be desired to keep the amount of lead chloride in the solution as low as possible, it is necessary to keep the concentration of the acid down to a minimum value. A certain amount of acid is usually required to keep other metals in solution; zinc, for example.

The concentration of the hydrogen sulphide in the solution is practically constant (0.0073 gram-molecules per litre at 20°) when the gas is passing through the solution. If the concentration of the hydrogen sulphide were large and the concentration of the metallic chloride small, a very large excess of acid would be needed to prevent metal being precipitated by the hydrogen sulphide. It will be observed, however, that the concentration of the hydrogen sulphide under ordinary circumstances is small. In consequence, a comparatively small amount of acid suffices to prevent the separation of sulphides of zinc, iron, nickel, cobalt, manganese. If the solubility of the hydrogen sulphide had been greater than it is some of the metals—zinc, iron, nickel . . . would have been included in the "hydrogen sulphide group"; and conversely, had the solubility of hydrogen sulphide been less than it is, some of the present members of the "hydrogen sulphide group" would not have been there. For instance, tin, lead, cadmium. . . .

The reaction with metal salt solutions and hydrogen sulphide is not always reversible. This is not because the equilibrium law has broken down, but rather because a consecutive reaction follows the precipitation

of the metal sulphide whereby the sulphide first precipitated changes into a different, maybe less soluble, variety. This may also explain the incomplete reversibility of the reaction between hydrogen sulphide and the zinc salts.

Under ordinary conditions, the solubilities of the sulphides in hydrochloric acid, starting with the least soluble, are approximately in the order:

As, Hg, Cu, Sb, Bi, Sn(ic), Cd, Pb, Sn(ous), Zn, Fe, Ni, Co, Mn.

Elements wide apart in the list can be easily separated by hydrogen sulphide in acid solutions, but elements close together in the list require a very careful adjustment of the amount of acid in solution before satisfactory separations can be made. For instance, the separation of cadmium or lead from zinc by means of hydrogen sulphide is only satisfactory when the concentration of the acid is very carefully adjusted. If too much acid be present, cadmium or lead will be imperfectly precipitated; while if too little acid be present, zinc will be precipitated with the cadmium or lead. Hence there is no sharp line of demarcation between metals precipitated and metals not precipitated by hydrogen sulphide from acid solutions. All depends upon the concentration of the acid. This is arbitrarily adjusted so that antimony, arsenic, lead, bismuth, cadmium, copper, mercury, and tin are precipitated by making the volume of the solution such that it contains approximately 10 c.c. of hydrochloric acid (sp. gr. 1.12) per 100 c.c. before passing the hydrogen sulphide. The aluminium, iron, zinc, nickel, cobalt, and manganese salts will be found in the filtrate. Barium, strontium, calcium, and magnesium salts will also be found in the filtrate along with the alkalis, because the sulphides of these elements are attacked and decomposed by water and by acids. *E.g.* $\text{Ca}_2\text{S} + 2\text{H}_2\text{O} \rightleftharpoons \text{Ca}(\text{OH})_2 + \text{Ca}(\text{SH})_2$.

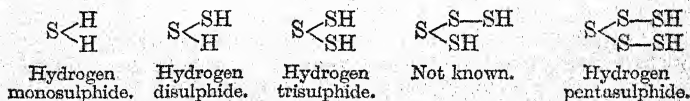
The above remarks can easily be translated into the language of ions. The precipitation is then supposed to proceed according to the equation: $\text{M}'' + 2\text{H}_2\text{S} \rightleftharpoons \text{M}(\text{HS})_2 + 2\text{H}^+$; or $\text{M}'' + \text{H}_2\text{S} \rightleftharpoons \text{MS} + 2\text{H}^+$. That is, the bivalent ion M'' reacts with the hydrogen sulphide, forming the sparingly soluble MS , or $\text{M}(\text{HS})_2$, which precipitates. In the process, hydrogen (acid) ions, H^+ , are formed. The hydrogen sulphide is itself supposed to be ionized in aqueous solution as indicated above: $\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}' \rightleftharpoons 2\text{H}^+ + \text{S}''$. The metal chloride, say, is also ionized: $\text{MCl}_2 = \text{M}'' + 2\text{Cl}'$. Hence the solution may be supposed to contain $\text{MCl}_2 + \text{H}_2\text{S} \rightleftharpoons \text{M}'' + 2\text{Cl}' + \text{S}'' + 2\text{H}^+$. When the solubility product $[\text{M}''] \times [\text{S}'']$ is exceeded, the solid MS separates from the solution, leaving hydrochloric acid ions behind: $2\text{H}^+ + 2\text{Cl}'$. A further amplification on the lines indicated in the text can now be made. Here, as elsewhere, it makes very little difference which mode of expression be used. The facts will stand for ever, and they are independent of fashion, taste, or caprice; the language used in describing the facts, like other customs, changes according to the prevailing or fashionable hypothesis, for, as T. Bergman long ago observed, fashion frequently determines arbitrarily what are to be considered as fictions, what as misrepresentations, and what as omissions.

§ 12. Hydrogen Persulphide, or Hydrogen Disulphide.

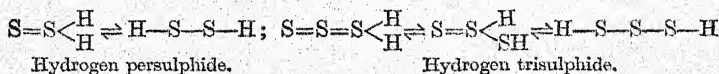
If dilute hydrochloric acid be poured into a solution of sodium polysulphide, say, Na_2S_5 , milk of sulphur is precipitated (p. 480): Na_2S_5

$+ 2\text{HCl} = 2\text{NaCl} + \text{H}_2\text{S} + 4\text{S}$. On the contrary, if the polysulphide be poured into the acid, little or no hydrogen sulphide is evolved, and a yellow oily liquid is obtained. This was once considered to be a mixture of hydrogen pentasulphide, H_2S_5 , with other hydrogen polysulphides. When the oil is distilled under reduced pressure, the fraction which is obtained at 69° under a pressure of 2 mm. of mercury is a pale yellow oil with a molecular weight, by the freezing-point method, corresponding with H_2S_3 —hydrogen trisulphide. The fraction which distils at 74° – 75° , under atmospheric pressure, has the composition H_2S_2 . This is hydrogen disulphide or hydrogen persulphide. The persulphide dissolves in benzene, forming a clear solution. Hydrogen persulphide is a colourless oily liquid with a specific gravity 1.376. It has a pungent irritating smell, and decomposes gradually into hydrogen sulphide and sulphur. The decomposition is faster in presence of water, and particularly alkalis. Mere contact with glass, paper, dust, etc., induces rapid decomposition. Hydrogen persulphide burns with a blue flame. Like its analogue, hydrogen peroxide, it has oxidizing and reducing qualities.

Hydrogen trisulphide, H_2S_3 , prepared as just described, resembles the persulphide in many of its properties, but its specific gravity is 1.496, and it solidifies between -52° and -54° . The three hydrogen sulphides— H_2S , H_2S_2 , and H_2S_3 —are all the hydrogen polysulphides whose individuality has been clearly demonstrated, although a whole ranging from H_2S_5 to H_2S_9 has been reported. There is a certain analogy between H_2O and H_2S ; and between H_2O_2 and H_2S_2 . The hydrosulphides are particularly interesting, for they bring out the peculiar property of sulphur atoms to form chains in which hydrogen can be replaced by an SH group:



These are sometimes called *chain formulæ*; the pentasulphide, for example, is virtually H-S-S-S-S-S-H . This is D. I. Mendeléeff's method of representing the constitution of the hydrogen sulphides. Several other schemes have been proposed, but no decisive evidence is available to justify one in preference to another. I. Bloch and F. Höhn (1908), indeed, assume that two or even three isomeric modifications may exist in equilibrium, side by side, at any given temperature:



It is assumed that at low temperatures Mendeléeff's chain type of molecules predominates; and at higher temperatures, when the colour darkens, that the more condensed type of molecules prevails.

§ 13. Sulphur Chlorides.

The study of the sulphur chlorides throws an interesting light on the valency of sulphur, and also on the constitution of some other sulphur compounds.

Sulphur chloride.—When dry chlorine is passed into sulphur heated in a retort, *A*, Fig. 171, the two elements combine directly, forming sulphur

chloride— S_2Cl_2 —which collects in the receiver *B*, cooled by a current of cold water, as a yellow liquid. The oil is purified by redistillation. The

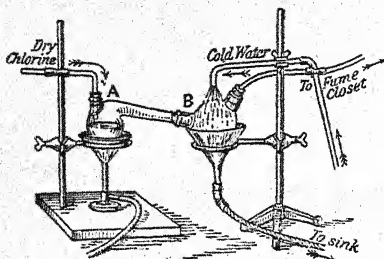


FIG. 171.—Preparation of Sulphur Chloride
— S_2Cl_2 .

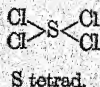
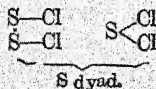
pale yellow liquid when pure has a pungent smell. It boils between 138° and 140° . The liquid fumes in moist air and is decomposed by water, forming sulphur dioxide, hydrochloric acid, and sulphur: $2S_2Cl_2 + 3H_2O = 4HCl + H_2SO_3 + 3S$. Sulphur chloride is used as a solvent for sulphur in the manufacture of rubber goods; for decomposing some minerals in analysis; and in preparing some anhydrous chlorides. Its vapour density is 67.5 ($H_2 = 1$),

corresponding with the molecule S_2Cl_2 , not SCl . Hence "sulphur monochloride" does not appear a very appropriate name for this compound. Sulphur chloride has the normal molecular weight in freezing benzene, and a partially polymerized molecule in boiling benzene or carbon tetrachloride. Selenium furnishes a similar selenium monochloride, Se_2Cl_2 .

Sulphur dichloride.—If chlorine be passed into sulphur chloride, cooled in ice, a garnet-red liquid, said to be SCl_2 , is obtained: $S_2Cl_2 + Cl_2 = 2SCl_2$. A mixture of liquid chlorine and sulphur chloride, S_2Cl_2 , in a sealed tube, at about 15° , gradually passes into a similar product. The molecular weight by the freezing point method in xylene, bromine, acetic acid, etc., corresponds with the molecule SCl_2 . It freezes at -80° ; and it can be distilled with but little decomposition under reduced pressures.

Sulphur tetrachloride.—If the temperature be still further reduced to -22° , still more chlorine is taken up by sulphur chloride, and sulphur tetrachloride, SCl_4 , is formed: $S_2Cl_2 + 3Cl_2 = 2SCl_4$. The red liquid begins to decompose into sulphur dichloride and chlorine directly it is removed from the freezing mixture at -22° . It freezes to a yellowish white solid between -30° and -31° . Sulphur tetrachloride unites with gold chloride and with stannic chloride, molecule for molecule, forming double compounds, e.g. $SnCl_4 \cdot SCl_4$.

Bromine forms sulphur bromide, S_2Br_2 , when heated with sulphur, but the existence of sulphur iodide is doubtful, because the product prepared by the evaporation of a mixed solution in carbon disulphide, though sold commercially as sulphur iodide, is considered to be a mixture or a solid solution. H. Moissan has isolated a sulphur hexafluoride— SF_6 —which is formed by the action of fluorine upon sulphur. The gas solidifies at -55° , and the liquid boils a few degrees higher. The gas is comparatively stable and inert chemically. These compounds of sulphur are interesting from the point of view of the valency hypothesis. Here is strong presumptive evidence of the variable valency of sulphur; and almost conclusive evidence that sulphur can be sexivalent:



It might be supposed that the properties of the sulphur fluorides would be analogous with those of the chlorides, but the resemblances are not very close. The fluorides are more stable than the chlorides.

Questions.

1. Describe the preparation of dry sulphuretted hydrogen from ferrous sulphide and outline experiments by which (a) hydrogen, (b) sulphur can be separated from it. Explain carefully why the molecular formula H_2S is assigned to sulphuretted hydrogen.—*Sheffield Univ.*

2. Give an account of the occurrence of the element of sulphur, and its compounds in nature. Describe the preparation and properties of the various modifications of sulphur.—*St. Andrews Univ.*

3. Explain fully the meaning of the symbolic equation : $FeS + 2HCl = FeCl_2 + H_2S$. If 6.8 grams of ferrous sulphide are taken, what would be the weight obtained of each of the substances on the right-hand side of the equation (Atomic weight of iron = 56; sulphur 32; chlorine 35.5)?—*Oxford Junior Locals.*

4. What do you know regarding the general behaviour of metallic sulphides towards (a) water, (b) hydrochloric acid, (c) caustic soda? Give equations.—*St. Andrews Univ.*

5. You are required to convert a given weight of sulphur into hydrogen sulphide as completely as possible. How would you proceed? What are the reactions of gaseous hydrogen sulphide respectively with (a) gaseous ammonia, (b) aqueous sodium hydroxide, (c) aqueous copper sulphate, and (d) gaseous sulphur dioxide?—*Univ. North Wales.*

6. If the heat of formation of HgO is 21.5 Cals. ; and of $HgO.HgCl_2$, 41 Cals. ; while if chlorine monoxide is an endothermal compound absorbing 18 Cals. when it is produced, how do you explain the formation of chlorine monoxide by the reaction $2Cl_2 + 2HgO = Hg_2OCl_2 + Cl_2O$, in harmony with the principle of maximum work?

7. How does sulphur occur in nature, and how is it obtained as stick sulphur and flowers of sulphur? Describe the changes which sulphur undergoes when heated, and give some account of its allotropic forms.—*New Zealand Univ.*

8. Crystals of native sulphur differ in their forms from crystals deposited from melted sulphur. Describe each form and name the system to which it belongs. The vapour density of sulphur is said to be abnormal at 500° and normal at 1000° . Explain the meaning of these statements.—*London Univ.*

9. What metallic solutions undergo the following changes when sulphuretted hydrogen is passed through acid solutions :—(a) First a reddish then a black precipitate, (b) first a white, then a dirty yellow, afterwards a black precipitate?—*Science and Art Dept.*

10. Cadmium yellow of commerce is a cadmium sulphide prepared by precipitating a cadmium salt with hydrogen sulphide. Cadmium was dissolved in dilute nitric acid so as to make a solution containing 21.2 per cent. $Cd(NO_3)_2$, how much cadmium and hydrogen sulphide are needed for treating 2 kilos. of solution? J. Bornemann, *Stoichiometrie*, Leipzig, 1909.

CHAPTER XXIII

COMPOUNDS OF SULPHUR WITH OXYGEN

§ 1. Sulphur Dioxide—Occurrence and Preparation.

Molecular weight, $\text{SO}_2 = 64.07$. Melting point, -73° ; boiling point, -8° . Relative density ($\text{H}_2 = 2$), 64.05; (air = 1) 2.264.

Occurrence.—Sulphur dioxide is found among the fumes from volcanic vents; in the springs of volcanic districts, and in the air of towns where it is derived from the sulphur compounds in the coal.

History.—The use of sulphur for disinfecting purposes has been known from very early times. It is referred to in Homer where Odysseus, after the slaughter of the suitors, and probably recognizing the need for a general cleansing, calls:

Quickly, O Nurse, bring fire that I may burn
Sulphur, the cure of ills.

J. Priestley (1770) prepared the gas by the action of hot concentrated sulphuric acid on mercury. Priestley called it sulphurous acid.

Preparation.—Sulphur dioxide is formed when sulphur burns in air: $\text{S} + \text{O}_2 = \text{SO}_2$. Between 6 and 8 per cent. of the sulphur is simultaneously oxidized to sulphur trioxide, SO_3 . If the sulphur be burnt in oxygen gas between 2 and 3 per cent. burns to sulphur trioxide.¹ The nitrogen in the air seems to favour the production of sulphur trioxide, whereas moisture and carbon dioxide do not affect the result appreciably. The proportion of sulphur trioxide formed increases with the pressure. When sulphur is oxidized by a peroxide—*e.g.* manganese peroxide—sulphur dioxide is formed: $\text{MnO}_2 + 2\text{S} = \text{MnS} + \text{SO}_2$. Sulphur dioxide is also formed when the sulphides of some metals are roasted in air—*e.g.* iron or copper pyrites: $4\text{FeS}_2 + 11\text{O}_2 = 2\text{Fe}_2\text{O}_3 + 8\text{SO}_2$. This reaction is very commonly used for making the sulphur dioxide employed in the manufacture of sulphuric acid.

The most convenient laboratory process, for small quantities, is to decompose commercial sodium bisulphite with concentrated sulphuric acid. A concentrated—40 per cent.—solution of sodium bisulphite, NaHSO_3 , is placed in a flask, which is then fitted with a tap funnel containing concentrated sulphuric acid, as indicated in Fig. 50. The gas can be washed by passing it through concentrated sulphuric acid. The reaction is symbolized: $\text{NaHSO}_3 + \text{H}_2\text{SO}_4 = \text{NaHSO}_4 + \text{H}_2\text{O} + \text{SO}_2$. The same gas

¹ The presence of sulphur trioxide in the sulphur dioxide formed when sulphur burns in air accounts for the "foggy" appearance of the gas. This will be understood when the properties of the trioxide have been studied.

is made by reducing sulphuric acid with charcoal, sulphur, silver, copper, or mercury. The charcoal process is used on a manufacturing scale when the presence of carbon dioxide, a by-product of the reaction, is not detrimental; for instance, in the manufacture of sulphites. The reaction is represented: $C + 2H_2SO_4 = 2H_2O + CO_2 + 2SO_2$. When sulphur is the reducing agent, the sulphur is simultaneously oxidized to sulphur dioxide: $S + 2H_2SO_4 = 2H_2O + 3SO_2$. In practice, the action is slow because the sulphur melts and offers but a small surface to the action of the acid. Copper and sulphuric acid are often employed in the laboratory when the bisulphite process is not convenient. Half fill a flask with copper turnings, and add sufficient sulphuric acid to not quite cover all the copper. The gas comes off when the flask is gently warmed. The apparatus is illustrated in Fig. 103, where two washing-bottles containing concentrated sulphuric acid are shown attached to the delivery tube in order to dry the gas.

The reaction is somewhat complex. It is usually symbolized: $Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2$. It is possible that the first action resembles the effect of zinc on dilute sulphuric acid whereby hydrogen is developed: $Cu + H_2SO_4 = CuSO_4 + H_2$. The hydrogen reduces the sulphuric acid to sulphur dioxide: $H_2 + H_2SO_4 = 2H_2O + SO_2$. In confirmation of this view hot concentrated sulphuric acid is reduced to sulphur dioxide by passing hydrogen through the liquid. Some of the sulphuric acid is possibly reduced to hydrogen sulphide. T. Cundall (1914) has shown that in addition to this reaction cuprous sulphate is probably one of the first products of the reaction at comparatively low temperatures: $8Cu + 4H_2SO_4 = 3Cu_2SO_4 + Cu_2S + 4H_2O$, and the cuprous sulphate reacts with sulphuric acid, forming cuprous sulphide and cupric sulphate: $5Cu_2SO_4 + 4H_2SO_4 = Cu_2S + 8CuSO_4 + 4H_2O$. The formation of cuprous sulphide can be shown by diluting the reddish-brown solution of copper in sulphuric acid and filtering through a Gooch's crucible. A red precipitate of copper is obtained: $Cu_2SO_4 = CuSO_4 + Cu$. The formation of cuprous sulphide is deduced from the fact that the reddish-brown liquor, on warming or standing, darkens and deposits a greyish-black precipitate of cuprous sulphide. There are a number of secondary reactions, since if the reddish-brown liquid be strongly heated, a greenish-black precipitate of cupric sulphide is deposited: $Cu_2S + 2H_2SO_4 = CuS + CuSO_4 + 2H_2O + SO_2$, and if heated with sulphuric acid, the cupric sulphide dissolves and is converted into the sulphate with the evolution of sulphur dioxide: $2CuS + 3H_2SO_4 = 2CuSO_4 + SO_2$. The production of hydrogen when sulphuric acid acts on copper is evidenced by the formation of aniline, $C_6H_5.NH_2$, when the action takes place at 130° in the presence of nitrobenzene, $C_6H_5.NO_2$, but this does not really prove that the hydrogen is even momentarily free.

§ 2. The Properties of Sulphur Dioxide.

Sulphur dioxide is a colourless gas with a smell characteristic of burning sulphur. Sulphur dioxide is an acute blood poison. According to Ogata (1884) 0.04 per cent. in air causes a difficulty in breathing after a few hours. Sulphur dioxide is also injurious to vegetation, and it is one of the "noxious vapours" complained about in manufacturing districts.

Aqueous solutions.—The gas is more than twice as heavy as air, and in consequence, it can be collected by the upward displacement of air. The gas cannot be collected satisfactorily over water because it is easily soluble in that menstruum. One volume of water at 0° dissolves 79.8 volumes of sulphur dioxide; and at 20° , 38.4 volumes. The aqueous solution is strongly acid, and it has the general properties characteristic of acids. It is hence called sulphurous acid, and is represented by the formula H_2SO_3 . The gas itself is accordingly called sulphurous anhydride. The dissolved gas is all expelled on boiling the aqueous solution. The existence of the compound H_2SO_3 is inferred from analogy with other acids, and from the general behaviour of aqueous solutions of the gas. The solution probably contains both dissolved sulphur dioxide and sulphurous acid. At 0° a crystalline hydrate— $\text{SO}_2 \cdot 6\text{H}_2\text{O}$, or $\text{H}_2\text{SO}_3 \cdot 5\text{H}_2\text{O}$ —is obtained; but several other crystalline hydrates have been reported. The aqueous solution readily decomposes into sulphuric acid and free sulphur when exposed to light: $3\text{H}_2\text{SO}_3 = \text{S} + 2\text{H}_2\text{SO}_4 + \text{H}_2\text{O}$.

The action of cold.—The gas is easily liquefied. A pressure of 1.5 atmospheres suffices for the condensation of the gas at 0° ; and at -10°

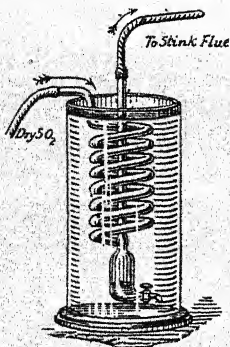


FIG. 172.—Gas Condensing Tube.

the gas liquefies under ordinary pressures. It is therefore sufficient to thoroughly dry the gas obtained by the action of copper on sulphuric acid by passing the gas through a couple of wash-bottles as indicated in Fig. 103, and then lead the gas through a condensing tube, Fig. 172, immersed in a freezing mixture: the freezing mixture is omitted in the drawing. The gas condenses to a clear, colourless, transparent, limpid liquid which boils at -8° and solidifies at -70° to a white snowlike mass. Liquid sulphur dioxide is sold commercially in thick glass "syphons," and where available, the "syphons" are used as a source of sulphur dioxide for laboratory work. By the evaporation of liquid sulphur dioxide a temperature approaching -50° can be obtained. Hence like ammonia and carbon dioxide, sulphur

dioxide is used as a refrigerating agent. Liquid sulphur dioxide is a good solvent for phosphorus, iodine, sulphur, resins, etc. The conductivity of these solutions is sometimes greater than that of aqueous solutions.

Dissociation of sulphur dioxide.—When the gas is passed through Deville's "hot and cold tube," Fig. 85, the surface of the silver tube is blackened owing to the formation of silver sulphide; at the same time, sulphur trioxide can be detected in the products of the reaction. Dry sulphur dioxide dissociates appreciably at 1200° , into sulphur and sulphur trioxide: $3\text{SO}_2 \rightleftharpoons \text{S} + 2\text{SO}_3$; and the moist gas into sulphur and sulphuric acid: H_2SO_4 . Like hydrogen sulphide, sulphur dioxide is decomposed by a series of electric sparks with the deposition of sulphur on the glass in the vicinity of the sparks. The reaction appears to be $3\text{SO}_2 \rightleftharpoons \text{S} + 2\text{SO}_3$. If a beam of light be sent through a long cylinder of the gas, at first, the gas appears to be clear and transparent, but in a few minutes the gas appears to decompose, for misty wavering striæ appear,

and gradually the whole tube appears to be filled with a fog. This action of light in certain gases is sometimes called Tyndall's effect. In the present case the effect appears to be due to the decomposition of the sulphur dioxide, probably $3\text{SO}_2 = 2\text{SO}_3 + \text{S}$. If left a short time in the dark the gas becomes clear owing to the recombination of the sulphur and sulphur trioxide, or to condensation of the solid products on the walls of the apparatus.

Oxidizing properties.—Sulphur dioxide is incombustible and a non-supporter of ordinary combustion. Some substances are able to burn in the gas by abstracting its oxygen. In this way, sulphur dioxide appears to act as an oxidizing agent. Ignited magnesium ribbon, for instance, continues to burn in the sulphur dioxide. Potassium, calcium, tin, lead, iron, arsenic, and antimony also burn when the finely-divided element is heated in an atmosphere of sulphur dioxide. The products of the reaction are oxides, sulphides, and occasionally free sulphur. Sulphur dioxide is decomposed by carbon at a high temperature: $2\text{SO}_2 + 3\text{C} = 2\text{CO} + \text{CO}_2 + 2\text{S}$. Sulphur is deposited when hydrogen sulphide and sulphur dioxide are brought into contact, say by placing a jar of sulphur dioxide and of hydrogen sulphide mouth to mouth. The gaseous exhalations from some volcanoes contain these two gases which, on mingling together, mutually decompose with the formation of sulphur (p. 472): $2\text{H}_2\text{S} + \text{SO}_2 = 2\text{H}_2\text{O} + 3\text{S}$. This reaction does not occur if the gases are thoroughly dried.

Freshly precipitated sulphide of iron, manganese, or zinc is readily dissolved by concentrated sulphurous acid, while sulphide of nickel or cobalt is dissolved with difficulty. The reaction has been recommended as a means of separating ferrous sulphide from titanium and zirconium hydroxide. The primary action of sulphurous acid on the sulphides in question is a double decomposition, $\text{MS} + \text{H}_2\text{SO}_3 = \text{MSO}_3 + \text{H}_2\text{S}$; the liberated hydrogen sulphide is oxidized by the sulphur dioxide, forming a thiosulphate, MS_2O_3 , thus: $2\text{MS} + 3\text{SO}_2 = 2\text{MS}_2\text{O}_3 + \text{S}$. No other sulpho-salts have been reported, although it is usually considered that dithionates, MS_2O_6 , are also produced. If the sulphurous acid be in great excess, acid sulphites, $\text{M}(\text{HSO}_3)_2$, may be formed.

Composition of sulphur dioxide.—When sulphur burns in oxygen, no change in volume occurs when the gas is cooled. The apparatus shown in Fig. 170 or Fig. 253 can be used. Taken in conjunction with Avogadro's hypothesis, the experiment proves that sulphur dioxide contains its own volume of oxygen; that one molecule of sulphur dioxide contains two atoms of oxygen, and that the formula of sulphur dioxide is S_nO_2 , where n is to be determined. Again, the relative density of sulphur dioxide is 64·046 ($\text{H}_2 = 2$). Hence:

One gram-molecule of sulphur dioxide weighs	64·05
One gram-molecule of oxygen weighs	32·00
Weight of sulphur in the molecule	32·05

The number 32·05 can represent one and only one atom of sulphur if 32·07 be the atomic weight of sulphur, since the two numbers are nearly the same.

The salts of sulphurous acid.—Two series of salts exemplified by KHSO_3 and K_2SO_3 are known; the former are the acid sulphites or the bisulphites, and the latter the normal sulphites. Hence sulphurous acid

is dibasic. With the exception of the alkaline salts, most of the sulphites are sparingly soluble in water. The alkaline sulphites are alkaline to litmus. If the sulphites are treated with strong acids, the anhydride, SO_2 , is evolved as indicated in one of the methods of preparation. The sulphites decompose on heating with the formation of sulphates: e.g. $4\text{Na}_2\text{SO}_3 \rightarrow 3\text{Na}_2\text{SO}_4 + \text{Na}_2\text{S}$; the bisulphites first lose H_2SO_3 before decomposing in this way: $2\text{NaHSO}_3 \rightarrow \text{Na}_2\text{SO}_3 + \text{H}_2\text{SO}_3$. The sulphites, as well as sulphurous acid itself, are readily oxidized. The presence of glycerine or sugar retards the rate of oxidation, possibly because the more viscous solutions offer some resistance to the free circulation of oxygen in the fluid. The monobasic acids HRSO_3 —where R is a radicle: CH_3 , C_6H_5 , etc.—are called sulphonie acids; and the monobasic acids of the type: HRSO_2 , sulphinic acids.

Reducing properties of sulphur dioxide.—Sulphur dioxide is a powerful reducing agent. It reduces permanganates to manganous salts; chromates to chromic salts; ferric to ferrous salts, etc. The latter reaction is often used in analytical work for the reduction of iron previous to its determination by volumetric processes.

The bleaching effects of sulphurous acid are due to its reducing properties. Moistened red rose-leaves, or fabrics dyed with, say, "magenta" dye, when placed in the gas lose their colour. The sulphur dioxide appears to react with the colouring matter, forming sulphuric acid and hydrogen: $\text{SO}_2 + 2\text{H}_2\text{O} = \text{H}_2\text{SO}_4 + \text{H}_2$, and the nascent hydrogen bleaches by reducing the colouring agent. This idea is supported by the fact that the colour of many articles bleached by sulphur dioxide can be restored by exposing the article to oxidizing conditions. The familiar yellow colour which gradually comes to bleached sponges, flannels, etc., is an example. The colour of bleached rose-leaves gradually returns when the rose-leaves are exposed to the air, or when the bleached leaves are dipped in dilute sulphuric acid. This shows that the colouring agent is not really destroyed during the bleaching. A coloured flower quickly drawn through the flame of burning hydrogen sulphide is bleached where the flame touches. This is not the case with burning sulphur—presumably owing to the absence of water. Dry sulphur dioxide, like dry chlorine, does not bleach.

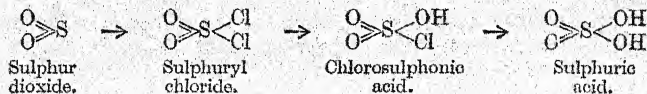
An aqueous solution of sulphur dioxide reduces chlorine, forming hydrochloric and sulphuric acids: $\text{SO}_2 + 2\text{H}_2\text{O} + \text{Cl}_2 = 2\text{HCl} + \text{H}_2\text{SO}_4$. Hence sulphurous acid is used as an "anti-chlor," that is, as an agent to remove the last traces of chlorine from articles bleached with chlorine. The goods are dipped in a solution of sulphur dioxide, or a sulphite. The extended use of the term reduction might here be emphasized. A reducing agent is a substance which can (1) remove oxygen from other substances; (2) decrease the non-metallic part of a compound, e.g. $\text{SnCl}_4 \rightarrow \text{SnCl}_2$; (3) add hydrogen to an element or compound, e.g. $\text{Cl} \rightarrow \text{HCl}$. The definition of an oxidizing agent—oxidation—can also be extended in the converse manner.

When sulphur dioxide is passed over lead peroxide, the reaction, $\text{PbO} + \text{SO}_2 = \text{PbSO}_4$, occurs, and the mass glows brightly. The sulphur dioxide here reduces the peroxide, so also sodium peroxide, when "dusted" into a cylinder of the gas, glows brightly owing to the heat evolved during the reaction: $\text{Na}_2\text{O}_2 + \text{SO}_2 = \text{Na}_2\text{SO}_4$.

Action on hydrogen iodide.—Sulphur dioxide reduces iodine to hydrogen iodide, HI, in the presence of water: $\text{SO}_2 + 2\text{H}_2\text{O} + \text{I}_2 = 2\text{HI} + \text{H}_2\text{SO}_4$. The reaction stops when a certain amount of hydrogen iodide has been formed—about 0.05 per cent. of acid. Again, as R. Bunsen showed in 1853, concentrated solutions of hydrogen iodide are oxidized by sulphuric acid, and the latter is reduced to sulphur dioxide: $\text{H}_2\text{SO}_4 + 2\text{HI} = \text{I}_2 + 2\text{H}_2\text{O} + \text{SO}_2$. This reaction is the reverse of that which immediately precedes. The two opposing reactions are in equilibrium when the speeds of the direct and reverse changes are equal. In other words, this reaction belongs to the type indicated on p. 119, and should be symbolized: $\text{I}_2 + \text{SO}_2 + 2\text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{SO}_4 + 2\text{HI}$. This reaction is important because the amount of sulphurous acid or of its salts in a given solution can be determined by adding a solution of iodine of known strength from a burette until the iodine solution is no longer decolorized. The equation furnishes the necessary data for the calculation. Every 253.84 grams of iodine corresponds with 64.07 grams of sulphur dioxide. The amount of sulphur dioxide must not exceed 0.05 per cent. or the "back reaction" will appreciably affect the results.

Action on iodic acid.—The reduction of iodic acid, HIO_3 , or of an iodate by contact with sulphur dioxide, results in the formation of sulphuric acid, and in the liberation of iodine: $2\text{HIO}_3 + 4\text{H}_2\text{O} + 5\text{SO}_2 = 5\text{H}_2\text{SO}_4 + \text{I}_2$. Strips of paper dipped in a solution of potassium iodate and starch turn blue in the presence of sulphur dioxide, see "Iodic acid."

Sulphuryl chloride.—If a mixture of sulphur dioxide and chlorine be exposed to direct sunlight—especially if a little camphor be present—a colourless liquid will be obtained which boils at 69° . The camphor acts as a catalytic agent (p. 160). The liquid is sulphuryl chloride— SO_2Cl_2 ; and the molecular weight in freezing benzene is normal SO_2Cl_2 . When treated with water, H—OH , the chlorine atoms in sulphuryl chloride can be replaced step by step by OH groups:



Chlorosulphonic acid, $\text{SO}_2\text{Cl(OH)}$.—This acid is best made by the direct union of sulphur trioxide and hydrogen chloride; or by distilling a mixture of concentrated sulphuric acid with phosphorus pentachloride, PCl_5 , or phosphorus oxychloride, POCl_3 , when: $2\text{H}_2\text{SO}_4 + \text{POCl}_3 = \text{HPO}_3 + \text{HCl} + 2\text{SO}_2\text{Cl(OH)}$. The liquid so obtained boils at 155.3° , and reacts with water with explosive violence, forming a mixture of sulphuric and hydrochloric acids. Sodium fluoride absorbs sulphur trioxide when it is exposed to the vapour from fuming sulphuric acid. The product when washed with alcohol gives a solution of sodium fluosulphonate, NaFSO_3 , and undissolved sodium fluoride. If the fluosulphonate is obtained by evaporating off the solvent, when the resulting solid is heated with fuming sulphuric acid, fluosulphonic acid, HFSO_3 , distils over. The fluosulphonates are fairly stable. Calcium fluosulphonate, $\text{Ca}(\text{FSO}_3)_2$, is formed by the action of fuming sulphuric acid on fluorspar. Fluosulphonic acid is stable at 900° , and does not decompose, as was once supposed, into sulphuric acid and sulphuryl fluoride, SO_2F_2 . When

boiled with sulphur, fluosulphonic acid furnishes hydrofluoric acid and sulphur dioxide: $2\text{HFSO}_3 + \text{S} = 3\text{SO}_2 + 2\text{HF}$.

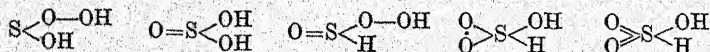
Uses.—Sulphur dioxide is used in the manufacture of sulphuric acid, as a refrigerating agent, as a solvent for extracting glue, gelatine, etc., for preserving meats, wines, etc. It prevents the growth of certain moulds, kills certain disease germs, etc. It is used for precipitating lime in sugar manufacture, as a bleaching agent for straw, silk, woollen, and goods too delicate for treatment with chlorine, in refrigerating machines (see p. 652), etc.

§ 3. The Constitution of Sulphurous Acid and the Sulphites.

More value is sometimes attached to a formula than that which it is intended to represent. In consequence of this, it has happened that a large number of chemists have regarded the determination of a formula for a compound as the great object to be accomplished, and they have forgotten that what we ought to know, and what is of vastly greater importance for the science, is the chemical conduct of the compound.—I. REMSEN.

The empirical formula of a compound is based upon its percentage composition, and the atomic hypothesis; while a molecular formula, in addition, is coupled with Avogadro's hypothesis. After the empirical and molecular formulæ have been determined, the chemical properties of the compound are studied in order to trace the relations between the atoms of the molecules, and when this has been done for all the atoms in the molecule of the substance, the constitution of the compound is said to have been determined. The result of the investigation is expressed as a structural or graphic formula. "It is assumed," as W. Lossen expressed it in 1880, "that the action of any particular atom on the other atoms in the molecule depends on the relative position of the atom in question; the properties and chemical behaviour of the molecule depend upon the actions of *all* the atoms on one another. Hence observations of the properties and the behaviour of a compound enable us to draw conclusions concerning the mutual actions of the atoms in the molecule of that substance, and the positions of the molecules relative to one another." A structural formula should summarize what is known about the chemical behaviour of the compound with respect to (1) the number and kind of atoms in the molecule; and (2) the relations between the atoms in the molecule.

There are several possible methods of representing the constitution of sulphurous acid, and accordingly of the sulphites. The sulphur may be bi-, quadri-, or sexivalent:



Sulphur dyad.

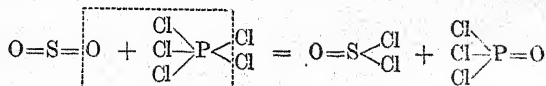
Sulphur tetrad.

Sulphur hexad.

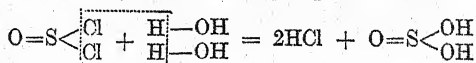
In one group of these formulæ, the hydrogen atoms are symmetrically placed with respect to the sulphur atom; and in the other group, the hydrogen atoms are unsymmetrical. Our problem is to select from these graphic formulæ the one which best represents the orientation of the atoms in the molecule of sulphurous acid.

Thionyl chloride.—If dry sulphur dioxide be treated with phosphorus pentachloride, PCl_5 , a colourless liquid—thionyl chloride, SOCl_2 —boiling

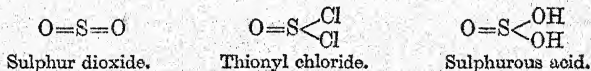
at 78°, is obtained. The initial and end products of the reaction are represented:



Sodium sulphite can be used: $\text{Na}_2\text{SO}_3 + 2\text{PCl}_5 = \text{SOCl}_2 + 2\text{NaCl} + 2\text{POCl}_3$. Thionyl chloride is also made by the action of chlorine monoxide upon sulphur at a low temperature, -12° , to prevent explosion. It is also made by adding sulphur trioxide to sulphur monochloride at 75° to 80° ; $\text{S}_2\text{Cl}_2 + \text{SO}_3 = \text{SOCl}_2 + \text{SO}_2 + \text{S}$. Thionyl chloride has the empirical formula SOCl_2 , and this also represents the molecular weight in freezing benzene; in boiling chloroform, the molecular weight has twice this value, being $(\text{SOCl}_2)_2$. Owing to its mode of formation from sulphur dioxide (sulphur quadrivalent), it is inferred that there is no change in valency of the sulphur atom when one of the oxygen atoms in sulphur dioxide is replaced by two chlorine atoms, and that the oxygen and two chlorine atoms are each directly attached to the sulphur. Thionyl chloride is decomposed by water, forming sulphurous acid.



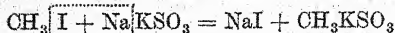
This reaction is interpreted to mean that sulphurous acid has a similar constitution to thionyl chloride; but the two chlorine atoms of the last-named compound have been replaced by two hydroxyl groups derived from the water. Hence it is inferred that the two hydroxyl groups of sulphurous acid are directly united with the sulphur atom. Otherwise expressed:



The constancy of structural arrangement.—It is commonly assumed that when an atom or group of atoms in a compound is replaced by another atom or group of atoms, the latter occupies the position vacated by the former without any essential structural change in the arrangement of the other atoms. This rule does not mean that the relations subsisting between the atoms of the molecule are not altered during the replacement of one atomic group by another, for the hydrogen atoms in, say, $\text{C}_2\text{H}_5\text{OCl}$ may be more or less easily affected by certain reagents than the H atoms in, say, $\text{C}_2\text{H}_5\text{OBr}$. The rule of the constancy of structural arrangement, and the action of water on thionyl chloride, make it probable that the two chlorine atoms of thionyl chloride are directly replaced by two hydroxyl groups; but let us inquire:

Does the molecule of sulphurous acid contain two hydroxyl groups symmetrically placed about the sulphur atom? If sulphurous acid contains its two hydroxyl groups symmetrically placed, we naturally prefer the formula $\text{HO}-\text{SO}-\text{OH}$. By neutralizing potassium hydrogen sulphite with sodium hydroxide, and by neutralizing sodium hydrogen sulphite with potassium hydroxide, two solutions are obtained from each of which crystalline potassium sodium sulphite can be separated. In both cases compounds with identical properties, namely, potassium

methyl sulphonate, CH_3KSO_3 , are obtained when the double sulphites are treated with methyl iodide, CH_3I . The reaction in each case is represented:



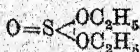
The methyl radicle CH_3 thus displaces the sodium atom, but not the potassium atom. Consequently, if two isomeric salts, say,



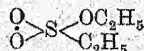
are capable of existing, the one in which the atom of potassium is directly connected with the sulphur atom is not stable and readily changes into the other. Otherwise expressed, there is at present no evidence of the existence of two different salts, $\text{KO}\cdot\text{SO}_2\cdot\text{Na}$ and $\text{NaO}\cdot\text{SO}_2\cdot\text{K}$. So far as we can tell, *the graphic formulæ of the sulphites correspond with two symmetrically placed hydroxyl groups in sulphurous acid*, and the formula is accordingly written: $\text{HO}-\text{SO}-\text{OH}$.

The "rule of the constancy of structural arrangement" might here easily lead us astray, because the fact that no isomerism has hitherto been detected in the compounds prepared by the two different methods may be due to the fact that with certain radicles only one configuration is stable and permanent. If another configuration be momentarily produced it immediately passes into the stable condition. Hence the negative results just obtained do not prove conclusively that an unsymmetrical sulphurous acid is non-existent, although the negative results may prove that there is only one stable configuration of NaKSO_3 or of CH_3KSO_3 under the conditions of the experiment.

Desmotropism or tautomerism.—Some compounds are known in organic chemistry which correspond with the existence of unsymmetrically placed hydroxyl groups in the sulphurous acid molecule. For instance, the action of ethyl alcohol on thionyl chloride, SOCl_2 , furnishes a compound $\text{SO}_3(\text{C}_2\text{H}_5)_2$. Since the structural formula of thionyl chloride is $\text{Cl}-\text{SO}-\text{Cl}$, the compound formed from it probably has the corresponding symmetrical structure $\text{SO}(\text{OC}_2\text{H}_5)_2$. It is accordingly called **symmetrical diethyl sulphite**. This salt boils between 158° and 158.5° , and its specific gravity is 1.1063 (0°). Again, the oxidation of ethyl mercaptan, $\text{C}_2\text{H}_5-\text{S}-\text{H}$ (in which the ethyl monad radicle, C_2H_5 , must be directly attached to the sulphur) furnishes a compound of the same ultimate composition, but with an unsymmetrical structure, hence, it is called **unsymmetrical diethyl sulphite**:



Symmetrical diethyl sulphite.



Unsymmetrical diethyl sulphite.

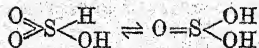
The unsymmetrical salt boils between 213° and 214° , and its specific gravity is 1.1712 (0°). The existence of these two sulphites does not necessarily mean that there are two different sulphurous acids; for there may be but one stable configuration of the acid itself under the conditions of the experiment.

It is not unlikely that under certain conditions, and by the action of certain reagents, the position of an hydrogen atom in the molecule

can change so that it behaves differently under one set of conditions from what it does under another set of conditions. As a result of the roving or wandering of the hydrogen atoms, a compound prepared by a chemical process may be different from what was anticipated though its composition be the same. The phenomenon appears an exception to the rule of the constancy of structural arrangement. As usual, the supposed phenomenon has been given a name, "tautomerism" (C. Laar, 1885)—from the Greek *ταυτό* (tauto), the same; *μέρος* (meros), a part—and also the alternative, perhaps better, name, "desmotropism" (P. Jacobson, 1887)—from the Greek *δεσμός* (desmos), a bond; *τρέπω* (trepein), to change. A substance is said to be tautomeric or desmotropic when it can react with other substances in such a way that it appears to be a compound with different constitutional formulæ.

N. Bland and J. F. Thorpe proposed to restrict the term *tautomerism* to those substances exhibiting the phenomenon under consideration in which one is not stable enough to be isolated in the free state; *desmotropism* to cases in which equilibrium mixtures of both forms are produced. The term *structural isomerism* is applied to cases where both forms are stable and can be isolated.

In any given system, the two desmotropic modifications of a substance are in a state of equilibrium which is so very sensitive to external influences that the one modification readily changes into the other. A desmotropic change is thus an intra-molecular phenomenon which only lends itself to observation under exceptional conditions. It is supposed that the position of a hydrogen atom in the molecule of sulphurous acid is labile, not rigid, because it can take up two different positions with respect to the other atoms. The change in the position of the hydrogen atom is accompanied by a change in the character of the linkages. In the present case, we have possibly:



We shall meet other examples later on—nitrous acid: $\text{H}-\text{NO}_2$ and $\text{HO}-\text{NO}$; hydrocyanic acid: $\text{H}-\text{C}:\text{N}$ and $\text{H}-\text{N}:\text{C}$; phosphorous acid; $\text{P}(\text{OH})_3$ and $\text{H.PO}(\text{OH})_2$, etc. The supposed desmotropism of the polysulphides has been already mentioned, p. 499.

§ 4. Hyposulphurous Acid and Hyposulphites.

C. L. Berthollet (1789) noticed that iron dissolves in sulphurous acid without giving off a gas; L. N. Vauquelin and A. F. Fourcroy (1798) found that tin and zinc behaved in a similar way; and C. F. Schönbein (1852) obtained some reactions with a lower sulphur acid, and which were probably due to what is now called hyposulphurous acid. P. Schützenberger discovered this acid in 1869.

Preparation.—By treating an aqueous solution of sulphur dioxide with finely divided zinc, a zinc salt of hyposulphurous acid is obtained: $2\text{H}_2\text{SO}_3 + \text{Zn} = \text{ZnS}_2\text{O}_4 + 2\text{H}_2\text{O}$. By using sodium hydrogen sulphite, the sodium salt is obtained. A yellow aqueous solution of the acid can be obtained by treating aqueous solutions of its salts with oxalic acid. It is probable that the zinc reacts with the sulphurous acid producing hydrogen: $\text{Zn} + \text{H}_2\text{SO}_3 = \text{ZnSO}_3 + \text{H}_2$; and that the nascent hydrogen

reduces the sulphurous acid to the acid in question: $2\text{H}_2\text{SO}_3 + 2\text{H} = \text{H}_2\text{S}_2\text{O}_4 + 2\text{H}_2\text{O}$. The acid then reacts with the zinc producing the zinc salt. The sodium salt has also been obtained in concentrated solution by the electrolysis of sodium hydrogen sulphite with a high density current. By adding sodium chloride to the solution, sodium hyposulphite, $\text{Na}_2\text{S}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$, is precipitated. This salt can be dehydrated at 60° . H. Moissan obtained the sodium salt by the action of sodium hydride on sulphur dioxide at about -40° , $4\text{NaH} + \text{H}_2\text{SO}_3 = \text{H}_2 + \text{Na}_2\text{S}_2\text{O}_4 + 2\text{NaOH}$. The potassium and lithium salts were obtained in a similar way; but with calcium and strontium hydrides, the reaction is complete only when the sulphur dioxide is under pressure.

Properties.—The acid is only known in aqueous solutions since it is rather unstable. It rapidly absorbs oxygen from the air, and is one of the most powerful reducing agents known. The sodium salt is used industrially, and in the chemical laboratory as a reducing agent. For instance, it bleaches sugar, indigo blue, etc. It reduces permanganates; and also reduces some metallic salts to the metal—e.g. silver nitrate, and gold and platinum chlorides; and in the case of copper sulphate, a brown copper hydride, Cu_2H_3 , is formed. It seems to act as an oxidizing agent towards hydrogen sulphide in the presence of moisture, for sulphur is precipitated, and thiosulphate is formed: $\text{Na}_2\text{S}_2\text{O}_4 + \text{H}_2\text{S} = \text{Na}_2\text{S}_2\text{O}_3 + \text{H}_2\text{O} + \text{S}$. The salts are called hyposulphites.¹

Composition.—The molecular weight of the acid has not been determined by the regular methods, but the molecular weight of the sodium salt, deduced from the depression of the freezing point of aqueous solutions, corresponds with the formula $\text{Na}_2\text{S}_2\text{O}_4$. Again, ammoniacal copper sulphate oxidizes a hyposulphite to a sulphite; iodine oxidizes it to a sulphate; and for every two atoms of sulphur present as hyposulphite, one atom of oxygen is required to oxidize it to sulphite: $\text{S}_2\text{O}_3 + \text{O} = 2\text{SO}_2$, or $2\text{Na}_2\text{S}_2\text{O}_4 + 2\text{H}_2\text{O} + \text{O}_2 = 4\text{NaHSO}_3$, and three atoms of oxygen to oxidize two atoms of sulphur to sulphate: $2\text{S}_2\text{O}_3 + 3\text{O}_2 = 4\text{SO}_3$. Hence it is inferred that hyposulphurous acid is derived from the anhydride S_2O_3 —sulphur sesquioxide—and not to the hypothetical *sulphoxylic anhydride*, SO , as supposed by P. Schützenberger. Schützenberger thought the acid had the formula H_2SO_3 , but A. Bernthsen's (1881) experiments, just cited, make it almost certain that the formula is $\text{H}_2\text{S}_2\text{O}_4$. The acid can be regarded as a mixed anhydride derived from an unsymmetric sulphurous acid, $\text{H.SO}_2\text{OH}$, and the hypothetical sulphoxylic acid, H.SO.OH , namely $\text{H.SO}_2\text{O.SO.H}$.

Sulphur sesquioxide, S_2O_3 , is supposed to be made as a malachite green crystalline mass by the direct union of sulphur and sulphur trioxide, and as a blue solution by the action of flowers of sulphur on Nordhausen sulphuric acid. There is some doubt whether the bluish-green mass is a mixture or a compound. With water, sulphur sesquioxide gives sulphur and a mixture of sulphuric and sulphurous acids, and not hyposulphurous acid. This does not prove that sulphur sesquioxide is not the anhydride of hyposulphurous acid because the latter decomposes in a similar way.

¹ The acid has also been called *hydrosulphurous acid*, but it is best to avoid this term. Do not confuse hyposulphurous acid with thiosulphuric acid, nor the salts with thiosulphates. Unfortunately, hyposulphurous acid is an old term for thiosulphuric acid, which has been abandoned by all but photographers.

§ 5. Sulphur Trioxide—Preparation.

Sulphur trioxide exists in at least two forms, α -sulphur trioxide and β -sulphur trioxide. Molecular weight, α - $\text{SO}_3 = 80.07$; β - $\text{SO}_3 = 160.14$. Melting point, α - SO_3 , 14.8° ; β - SO_3 , 25° ; boiling point of both varieties, 46.2° .

History.—Sulphur trioxide seems to have been made by B. Valentine in the fifteenth century, and called "philosophical salt." K. W. Scheele (1777) and G. de Morveau (1786) called it the anhydride of sulphuric acid.

Preparation.—Sulphur trioxide is made by distilling ferric sulphate: $\text{Fe}_2(\text{SO}_4)_3 = \text{Fe}_2\text{O}_3 + 3\text{SO}_3$; fuming sulphuric acid; or sodium pyrosulphate: $\text{Na}_2\text{S}_2\text{O}_7 = \text{Na}_2\text{SO}_4 + \text{SO}_3$; and also by heating concentrated sulphuric acid with phosphorus pentoxide when $\text{H}_2\text{SO}_4 + \text{P}_2\text{O}_5 = \text{SO}_3 + 2\text{HPO}_3$. In the latter reaction the phosphorus pentoxide is able to dehydrate, that is, remove the elements of water from sulphuric acid.

As indicated on p. 502, a small amount of sulphur trioxide is formed

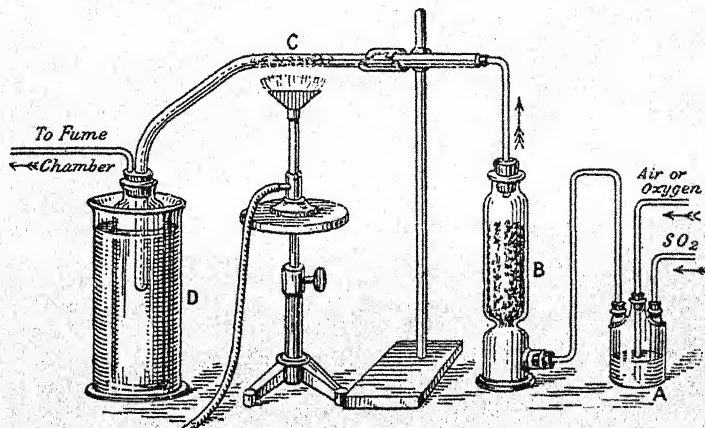


Fig. 173.—Preparation of Sulphur Trioxide (Contact Process).

when sulphur burns in air, or in oxygen. Sulphur dioxide is scarcely affected when warmed with oxygen, but if a mixture of sulphur dioxide and oxygen be passed over platinized asbestos at about 400° , the oxidation of the sulphur dioxide is nearly complete. The platinized asbestos acts as a catalytic agent (p. 124). Instead of platinized asbestos, ferric oxide, vanadium oxide, copper oxide, chromic oxide, and pumice stone, etc., have been employed, but none are so effective as platinized asbestos. With ferric oxide, for instance, the yield of sulphur trioxide is not much over 60 per cent. of the theoretical, under conditions where it is approximately complete with platinized asbestos. The action can be illustrated by an apparatus similar to that depicted in Fig. 173. The sulphur dioxide and oxygen are sent, from gas-holders, through a wash-bottle, A, shown on the right of the diagram, containing concentrated sulphuric acid. The mixed gases travel through a tower, B, containing pumice stone saturated with concentrated sulphuric acid. The dried mixture of oxygen and

sulphur dioxide then passes over platinized asbestos, *C*, warmed to about 400°, in a hard glass tube. Combination occurs, and the sulphur trioxide condenses in the tube and bottle, *D*, surrounded by a jar containing a freezing mixture, and shown on the left of the diagram.

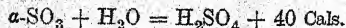
§ 6. The Properties of Sulphur Trioxide.

In 1876 R. Weber showed that sulphur trioxide exists in two modifications—solid and liquid at ordinary temperatures.

***α*-Sulphur trioxide—liquid sulphur trioxide.**—This is a colourless liquid, boiling at about 44·88°, obtained by repeated distillation of the product of the reactions just mentioned. The liquid fumes strongly in air. It crystallizes in long prismatic needles between 16° and 17°, and melts at 16·8°; the specific gravity at 20° is 1·9229. The vapour density (air = 1) varies from 2·74 to 2·76. These numbers and the lowering of the freezing point of phosphorus oxychloride correspond with a molecular weight of 80.

***β*-Sulphur trioxide—solid sulphur trioxide.**—If *α*-sulphur trioxide be maintained at 16° for some time, it gradually passes into solid crystals which look like asbestos. The mass fumes in air, and gradually passes into the *α*-form if kept at a temperature between 50° and 100°. The molecular weight by the depression of the freezing point of phosphorus oxychloride corresponds with 160, the theoretical value for S_2O_6 .

The two forms are not very different chemically although the *β*-sulphur trioxide reacts somewhat less vigorously than the *α*-variety. Both dissolve in water with a hissing sound as if a piece of red-hot iron were plunged into the water; much heat is evolved, and sulphuric acid is produced:



Hence sulphur trioxide is also called sulphuric anhydride. The anhydride does not redden blue litmus if extreme precautions be taken to exclude moisture. Sulphur trioxide reacts directly with many metallic oxides, forming the corresponding sulphates, *e.g.* with barium oxide, $\text{BaO} + \text{SO}_3 = \text{BaSO}_4$; so much heat is evolved during the action that the mass becomes incandescent.

The solid trioxide can be vaporized without melting. At low temperatures the vapour seems to contain both SO_3 and S_2O_6 molecules in equilibrium: $\text{S}_2\text{O}_6 \rightleftharpoons 2\text{SO}_3$. The dissociation is completed as the temperature rises. When heated to 1000°, sulphur trioxide decomposes completely into sulphur dioxide and oxygen: $2\text{SO}_3 = 2\text{SO}_2 + \text{O}_2$. Two volumes of sulphur trioxide produce two volumes of sulphur dioxide and one volume of oxygen. This corresponds with the formula $(\text{SO}_3)_n$. The molecular weight of the *α*-form, indicated above, shows that its formula must be SO_3 ; and the *β*-form, S_2O_6 . The graphic formula for SO_3 may be



Sulphur sexivalent.



Sulphur quadrivalent.



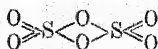
Sulphur bivalent.

The first formula is usually taken in harmony with the supposed sexivalency

of sulphur in sulphur hexafluoride, and it fits in best with facts to be developed later. The graphic formula for the two varieties will then be



α -Sulphur trioxide.



β -Sulphur trioxide or sulphur hexoxide.

The thermal values of the reactions, $\text{S} + \text{O}_2 = \text{SO}_2 + 69.26$ Cals., and of $\text{SO}_2 + \text{O} = \text{SO}_3 + 22.73$ Cals., are worthy of note. It follows that the thermal value of each of the first two oxygen atoms is 34.63 Cals., and the thermal value of the next atom of oxygen is 22.73 Cals. Hence it has been said that the third atom of oxygen is "less firmly attached" to the sulphur atom than the other two; and further that the six valencies of sulphur are not equivalent. This is supposed to correspond with the fact that sulphur trioxide is very readily dissociated into sulphur dioxide and oxygen. Some consider sulphur trioxide to be an "oxide of sulphur dioxide." The inference is inconclusive and faulty.

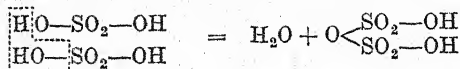
Occurrence.—Solutions of sulphur dioxide soon absorb oxygen from the air, forming sulphuric acid, hence sulphuric acid may be found in water from sulphurous springs in volcanic districts; *e.g.* the water of the Rio Vinagre (Mexico) contains 0.11 per cent. of H_2SO_4 calculated as SO_3 , and also 0.09 per cent. of HCl. Pit water and streams of water which have been in contact with oxidizing pyrites may also contain free sulphuric acid. This is often a source of trouble industrially owing to the corrosive action of such water on pump valves, etc. The occurrence of combined sulphuric acid as sulphates was indicated on p. 472.

§ 7. Pyrosulphuric Acid and Pyrosulphates.

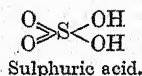
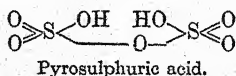
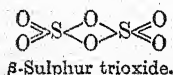
Sulphur trioxide dissolves in concentrated sulphuric acid. The solution is often called Nordhausen or fuming sulphuric acid. It is made by distilling dehydrated ferrous sulphate, $\text{FeSO}_4 \cdot \text{H}_2\text{O}$ (preferably with a little sulphuric acid or water in the receiver). The reaction appears to take place in two stages: $6\text{FeSO}_4 \cdot \text{H}_2\text{O} = \text{Fe}_3(\text{SO}_4)_3 + 2\text{Fe}_2\text{O}_3 + 3\text{SO}_2 + 6\text{H}_2\text{O}$; followed by $\text{Fe}_2(\text{SO}_4)_3 = \text{Fe}_2\text{O}_3 + 3\text{SO}_3$. The same acid is obtained by heating sodium pyrosulphate admixed with concentrated sulphuric acid: $\text{Na}_2\text{S}_2\text{O}_7 + \text{H}_2\text{SO}_4 = 2\text{NaHSO}_4 + \text{SO}_3$. Sodium pyrosulphate is made by heating sodium bisulphate to about 400° : $2\text{NaHSO}_4 = \text{H}_2\text{O} + \text{Na}_2\text{S}_2\text{O}_7$. Hence the term "pyro"—from the Greek $\pi\rho\upsilon$ (pyr), fire. This salt is sometimes called "anhydrous sodium bisulphate." The sodium bisulphate formed in the manufacture of fuming sulphuric acid is converted back to the pyrosulphate by heating it to about 400° .

Fuming sulphuric acid is a viscous, oily-looking liquid which is considered to be a solution of variable proportions of sulphur trioxide in concentrated sulphuric acid. The "fuming" of the acid is due to the escape of the sulphur trioxide. If the fuming acid be warmed, sulphur trioxide volatilizes and leaves sulphuric acid, H_2SO_4 , behind. This acid has often a brownish colour owing to the presence of a little organic matter. It may also contain other impurities. Fuming sulphuric acid is used in refining petroleum, in the manufacture of dyes, explosives, shoe-blackening, etc. When cooled below 0° , the fuming acid furnishes crystals of pyrosulphuric acid, $\text{H}_2\text{S}_2\text{O}_7$ or $\text{H}_2\text{SO}_4 \cdot \text{SO}_3$ or $\text{H}_2\text{O} \cdot 2\text{SO}_3$, melting at 35° .

The last mode of expressing the composition shows the origin of the term disulphates sometimes applied to the pyrosulphates. Two molecules of sulphuric appear to be condensed into one molecule of disulphuric acid with the elimination of one molecule of water: $2\text{H}_2\text{SO}_4 = \text{H}_2\text{O} + \text{H}_2\text{S}_2\text{O}_7$. The salts are dibasic. The bivalent acid radicle is " S_2O_7 ." The graphic formula is indicated in the scheme:



on the assumption that the formula for sulphuric acid is $\text{SO}_2(\text{OH})_2$; that disulphuric or pyrosulphuric acid is formed by the abstraction of a molecule of water from two molecules of sulphuric acid; and that when an anhydride is formed by the abstraction of water from a molecule of acid, with each molecule of water abstracted from the acid, one oxygen atom takes the place previously occupied by two hydroxyl groups. The relation between pyrosulphuric acid, sulphur hexoxide, and sulphuric acid is shown by the following schemes:



Hence sulphur hexoxide, or β -sulphur trioxide may be regarded as the anhydride of pyrosulphuric acid; and sulphur trioxide, or α -sulphur trioxide, is the anhydride of sulphuric acid. Sulphur trioxide and sulphur chloride, S_2Cl_2 , react, forming sulphur dioxide and pyrosulphuryl or disulphuryl chloride, $\text{S}_2\text{Cl}_2 + 5\text{SO}_3 = 5\text{SO}_2 + \text{S}_2\text{O}_5\text{Cl}_2$. This chloride is a colourless fuming liquid which boils at about 151° (730 mm.) and reacts with water to form sulphuric and hydrochloric acids. The vapour density at 205° agrees with $\text{S}_2\text{O}_5\text{Cl}_2$ if no moisture be present. In pyrosulphuryl chloride, $\text{Cl}-\text{SO}_2-\text{O}-\text{SO}_2-\text{Cl}$, the two hydroxyl radicles of pyrosulphuric acid, $\text{HO}-\text{SO}_2-\text{O}-\text{SO}_2-\text{OH}$, are replaced by chlorine.

Consonant with the definitions of an acid and of an anhydride, pyrosulphuric acid appears at the same time to be an acid and an anhydride. The three phosphoric acids, to be studied later, form a more striking illustration of this idea.

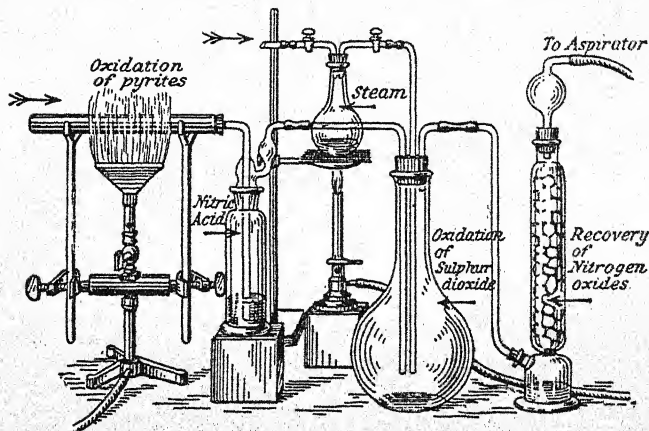
§ 8. The Manufacture of Sulphuric Acid—Chamber Process.

Sulphuric acid plays so important a rôle in commercial chemistry that it has been said that a country's prosperity can be estimated from its output of this acid.—ANON.

Sulphuric acid is one of the most important of acids. It is used in most of the chemical industries. Something like 4,000,000 tons are consumed yearly. The acid is used for making sodium carbonate, hydrochloric acid, nitric acid, etc.; it is used in the manufacture of explosives, fertilizers, alum, glucose, phosphorus, dyes, etc.; and it is employed for various operations in dyeing, bleaching, electroplating, metallurgy, etc. A carboy of the commercial acid (sp. gr. 1.84) holds about 180 lbs. and the commercial acid sells at 9s. per cwt. retail and the pure acid at about $2\frac{3}{4}d$.

per lb. Good colourless commercial acid contains 95 to 96 per cent. of H_2SO_4 . The so-called *chamber acid* has about 50 to 53 per cent. of SO_3 , or 62 to 65 per cent. H_2SO_4 ; *oil of vitriol* has about 75 to 78 per cent. of SO_3 , or 92 to 95 per cent. H_2SO_4 ; and the *fuming acid* or *oleum* has about 30 per cent. free SO_3 , or 87.14 per cent. free and combined SO_3 . The impurities in the commercial acid are nitric acid, nitrogen oxides, lead sulphate, arsenic, iron, selenium, calcium sulphate, haloid acids, etc.

Sulphuric acid can be prepared on the large scale by the decomposition of certain sulphates—as just indicated in connection with fuming sulphuric acid—or by burning sulphur or sulphides to sulphur dioxide, and oxidizing the latter to sulphur trioxide in the presence of moisture. The oxidation of sulphur dioxide under the influence of moisture alone is very slow, but in the presence of certain catalytic agents the oxidation proceeds much



Burners → Glover's tower → Lead chambers → Gay-Lussac's tower.

Fig. 174.—Illustrative Model of the Chamber Process

more quickly. Accordingly, the dry gases are passed over platinized asbestos or other solid catalytic agent (contact process); or mixed with nitrogen oxides in large chambers (chamber process).

Some of the principles involved in the manufacture of sulphuric acid by the lead chamber process are illustrated by the laboratory apparatus, Fig. 174. The sulphur or iron pyrites is heated in a porcelain boat placed in a quartz or porcelain tube. The mixture of air and sulphur dioxide is passed through a wash-bottle containing nitric acid, and thence into a large flask. The large flask is represented by large "lead chambers" on a real factory. Steam is also generated in a flask connected with the same globe. The amount of air drawn in with the steam is regulated by means of the stopcocks. The sulphur dioxide is oxidized to the trioxide in the presence of the nitrogen oxides which were mixed with the gases as they bubbled through the nitric acid. The steam reacts at once with the trioxide, etc., forming sulphuric acid which collects at the bottom of the flask. The nitrogen oxides pass along with the current of air, etc., to

the base of the tower on the right of the diagram. The tower is filled with coke soaked in concentrated sulphuric acid. This is represented by the Gay-Lussac's tower on a large works. The red fumes visible at the base of the tower are absorbed readily higher up, and the globe is placed at the top of the tower to enable this fact to be demonstrated. The flask or tower is connected with an aspirator which draws gases through the whole system. The aspirator is represented by a chimney on a real works.

If insufficient steam be associated with the reacting gases—*i.e.* if the current of steam is stopped by turning the necessary stopcock—only sulphur dioxide, air, and nitrous fumes pass into the glass globe. In that case, crystals of nitrosulphonic acid, also called nitrosylsulphuric acid— $\text{SO}_2\text{OH}\cdot\text{NO}_2$ or $\text{SO}_2\text{OH}\cdot\text{NO}$ —the so-called chamber crystals, condense on the walls of the vessel. The crystals decompose with the evolution of red fumes when steam is passed into the flask.

How does the catalytic agent work?—According to an old theory of E. M. Péligot (1844), nitric oxide, NO , unites with the oxygen from the air, forming nitrogen peroxide: $2\text{NO} + \text{O}_2 = 2\text{NO}_2$. The nitrogen peroxide then oxidizes the sulphur dioxide, and is at the same time reduced to nitric oxide: $\text{NO}_2 + \text{SO}_2 + \text{H}_2\text{O} = \text{NO} + \text{H}_2\text{SO}_4$. The nitric oxide is again oxidized to the peroxide, and so the cycle commences anew.¹ It is generally agreed that before the sulphuric acid is formed, the sulphur dioxide actually unites with the nitrogen oxide—say, nitrous acid—to form an “intermediate compound.” It is inferred that under normal conditions the intermediate compound decomposes as fast as it is formed. This inference appears to have been made because the elusive intermediate compound has not been isolated (p. 163). At present, therefore, the known facts accord best with Péligot's simple hypothesis. Speaking generally, whenever a question cannot be answered by an appeal to facts in the laboratory, keen, sometimes bitter, controversies are waged over schemes devised in the study. In the present case, quite an array of *imaginary intermediate compounds* have been advocated— $\text{H}_2\text{N}_2\text{SO}_6$, E. Divers, 1911; H_2NSO_5 , F. Raschig, 1911; HNSO_5 , H. Davy, 1812; etc. Taking E. Diver's scheme as illustrative of current theories, the intermediate compound nitroxysulphuric acid, $\text{H}_2\text{N}_2\text{SO}_6$, is supposed to be formed by direct combination of nitrous acid with sulphur dioxide: $2\text{HNO}_2 + \text{SO}_2 = \text{H}_2\text{N}_2\text{SO}_6$, and the nitroxysulphuric acid so formed is immediately decomposed into nitric oxide and sulphuric acid: $\text{H}_2\text{N}_2\text{SO}_6 = 2\text{NO} + \text{H}_2\text{SO}_4$. The nitric oxide so formed is immediately oxidized to nitrogen peroxide: $2\text{NO} + \text{O}_2 = \text{N}_2\text{O}_4$. This again unites with sulphur dioxide and water to reform nitroxysulphuric acid and oxygen, $2\text{N}_2\text{O}_4 + 2\text{H}_2\text{SO}_4 = 2\text{H}_2\text{N}_2\text{SO}_6 + \text{O}_2$. The nitroxysulphuric acid decomposes as indicated above and the cycle begins anew. It must be added that the mere fact that a definite compound can be isolated by altering the conditions under which a reaction progresses cannot be accepted as a conclusive proof that the compound so isolated is necessarily an intermediate compound when the reaction progresses under other conditions. A great

¹ J. J. Berzelius had a similar theory in 1835 in which nitrogen trioxide— N_2O_3 —was supposed to be formed instead of nitrogen peroxide— NO_2 . The objection to this intermediate compound, not necessarily valid, will be obvious from the discussion on nitrogen trioxide.

number of definite compounds have been prepared by the mutual action of the nitrogen and sulphur oxides with and without water.

The manufacturing process.—Sulphuric acid was known to Geber and other writers in the thirteenth century. B. Valentine, in the fifteenth century,¹ described the preparation of sulphuric acid from calcined ferrous sulphate and silica; and by burning sulphur and nitre in moist air. A. Sala made it in 1630 by burning sulphur in the presence of steam and an excess of air; and N. Lefèvre and N. Lemery improved Sala's method by using nitre. A. Libavius (1595) recognized the identity of sulphuric acid made by the different processes; previously, the acids were supposed to be different. Ward, in 1740, prepared sulphuric acid by burning sulphur with nitre in large glass vessels—40 to 60 gallons capacity—time after time, until the acid which collected on the bottom of the vessels was strong enough to pay for its concentration in glass retorts. The acid was sold as "oil of vitriol made by the bell," to distinguish it from the acid made from ferrous sulphate. Ward's process reduced the price of sulphuric acid from 2*s.* 6*d.* per ounce to 2*s.* per pound. Roebuck and Garbett

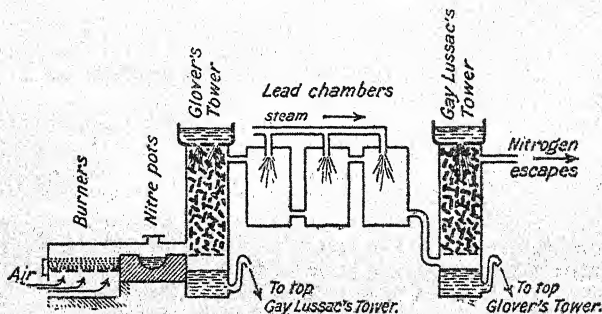


FIG. 175.—Diagrammatic Illustration of the Manufacture of Sulphuric Acid by Chamber Process.

substituted lead chambers for Ward's glass vessels in a works at Birmingham in 1746. F. Clément and J. B. Désormes, in 1793, showed that the process could be made continuous, and that the nitre plays an intermediary part between the sulphur dioxide and the air. J. L. Gay-Lussac's tower was invented in 1827; and J. Glover's tower was used at Newcastle-on-Tyne in 1859. The general disposition of the manufacturing plant is indicated by the diagrams, Fig. 175.

1. The burners.—The sulphur dioxide is made in the burners, in various ways, *e.g.* (1) by burning sulphur (brimstone); (2) by heating pyrites, etc., in a current of air; (3) by burning hydrogen sulphide, etc.; and (4) from the flue gases produced in smelting ores—thus the Tennessee Copper Co. is said to make 200,000 tons of acid per annum in this way, while the Anaconda Co. in 1912 were said to lose the equivalent of $1\frac{1}{2}$ million tons of acid per annum *via* the flue gases at their Washoe works. Air is drawn through the burners in excess of that required for oxidizing the sulphur. The necessary draught is regulated by chimney, etc.

¹ As in the case of Geber, some writings attributed to Valentine are supposed to be forgeries.

2. The Glover's tower.—The hot mixture of air and sulphur dioxide passes up a tower packed with flints down which trickles a mixture of a weak acid from the lead chambers and the strong nitrated acid which has been used to absorb nitrous fumes, and recovered in the Gay-Lussac's tower to be described later. The functions of the Glover's tower are: (1) recover the nitrous oxides from Gay-Lussac's tower; (2) cool the gases from the burners; (3) help to concentrate the acid trickling down the tower; (4) partly oxidize the sulphur dioxide from the burners; and (5) introduce the necessary nitric acid into the lead chambers by running nitric acid down the tower along with the nitrated acid from the Gay-Lussac's tower.¹ Before the acid reaches the foot of the tower it is fully denitrated.

3. The lead chambers.—The mixture of air, sulphur dioxide, and nitrous fumes passes into a set of three leaden chambers—may be of a total capacity of 150,000 cubic feet—into which steam is blown from low-pressure boilers. The gases in the chambers are thus intimately mixed. The oxidation of the sulphur dioxide mainly occurs in the first two chambers. The gases are passed through the chambers slowly so as to allow time for all the sulphur dioxide to be oxidized. The third chamber serves mainly to dry the gases. The chambers are kept cool enough to condense the sulphuric acid which collects on the floor and is drawn off periodically. It is called *chamber acid*. Chamber acid contains between 62 and 70 per cent. H_2SO_4 .

4. The Gay-Lussac's tower.—The excess air which leaves the lead chambers is highly charged with nitrogen oxides. These are recovered by causing the exit gases from the chambers to pass up a tower, packed with coke, down which concentrated sulphuric acid is trickling. The concentrated acid absorbs the nitrous fumes. The "nitrated acid" which collects at the foot of the Gay-Lussac's tower is pumped to the top of the Glover's tower along with some of the more dilute chamber acid. The "nitrated acid" trickling down the Glover's tower loses the absorbed nitrous fumes and some water. The acid which collects at the foot of the tower contains about 80 per cent. H_2SO_4 .

5. The concentration of the acid.—The chamber acid is not allowed to attain a higher strength than about 62 to 70 per cent. H_2SO_4 , because a stronger acid begins to absorb the nitrous fumes from the chambers. The chamber acid is therefore concentrated further either in the Glover's tower as indicated above, or in leaden concentrating pans until it contains about 79 per cent. of H_2SO_4 . If the acid be much stronger than this, it begins to attack the lead evaporating pans rather seriously. This acid is now called "B.O.V." ("brown oil of vitriol"), its usual brown colour is due to the presence of organic matter. The further concentration of the acid is effected either by the boiling it in glass, vitrified quartz, or in platinum stills. The acid may also be placed in these stills direct from the chambers. At first, very weak sulphuric acid distils over. The concentration of the acid in the still gradually rises until it has about 98.3 per cent. H_2SO_4 . Any further concentration cannot be done by evaporation since the acid itself then distils over. The further concentration of the acid, in the rare event of its being required, is effected by cooling the concentrated

¹ The gases from the Glover's tower sometimes meet a pipe bringing in nitrous fumes made by heating concentrated sulphuric acid with Chili saltpetre by the waste heat from the burners.

acid when crystals of 100 per cent. H_2SO_4 , melting at 10.5° , separate: the acid can of course be further "concentrated" by dissolving sulphur trioxide in it—see "Fuming sulphuric acid."

§ 9. The Manufacture of Sulphuric Acid—Contact Process.

The manufacture of sulphuric acid by the contact process can be illustrated by the apparatus depicted in Fig. 173, if the jar, etc., *D*, on the left, be replaced by a tower, packed with coke, Fig. 176, down which water or dilute sulphuric acid is allowed to trickle. The liquid absorbs the sulphur trioxide. The oxidation of the sulphur dioxide really involves two opposing reactions represented by the equation: $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$. By passing the mixed gases—air and sulphur dioxide—through a porcelain tube in the absence of platinized asbestos, about 10 per cent. of sulphur trioxide is formed at 400° under the conditions of manufacture, whereas in the presence of platinized asbestos, very nearly 100 per cent. is oxidized.

Without the platinized asbestos, a maximum 30 per cent. oxidation is obtained just over 600° . If the temperature, with platinized asbestos, is above or below 400° , the yield of sulphur trioxide is reduced; for instance, at 250° there is a 45 per cent. conversion, and at 1000° no

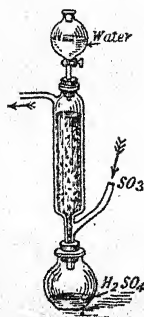


FIG. 176.—Condenser for SO_3 .

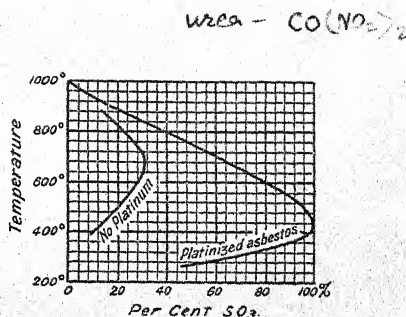


FIG. 177.—Effect of Temperature on the Oxidation of Sulphur Dioxide

sulphur dioxide is oxidized at all. The curves, Fig. 177, show the effect of temperature on the yield of sulphur trioxide. The presence or absence of the catalytic agent does not affect the final state of true equilibrium, but it does affect the speed at which the equilibrium is attained. In consequence, the final state of equilibrium appears different with and without the platinized asbestos because the process requires an indefinite time to attain a state of equilibrium without the catalytic agent. The curves, Fig. 177, thus represent the per cent. of sulphur trioxide formed when the mixed gases are passed slowly through a tube with and without platinized asbestos at the temperatures stated. Not only is the speed of oxidation of the sulphur dioxide determined by the temperature, but it is accelerated by increasing the proportion of oxygen beyond that represented by the above equation: two volumes of sulphur dioxide per one volume of oxygen. Experience shows that the most efficient mixture contains three molecules of oxygen per two molecules of sulphur dioxide. The use of ferric oxide in place of platinum as a catalyst—in the so-called Mannheim-Clemm system—has been previously discussed. Ferric oxide has some advantages over platinum, for instance, its cheapness permits it to be replaced

frequently, and it is not so sensitive to poisonous impurities—*e.g.* arsenic, indeed, it actually absorbs arsenic and so acts as a purifying agent. One disadvantage is that it requires a higher temperature (about 700°) to do its work, and the oxidation is incomplete (50 to 60 per cent.); hence ferric oxide can be used in conjunction with platinum, first oxidizing the mixed gases at a high temperature with the ferric oxide catalyst, and absorbing the resulting sulphur trioxide, then passing the remaining gases over platinum at a lower temperature so as to complete the reaction.

The manufacturing process.—W. Ostwald has pointed out that if time is money, then catalysis, which may be said to accelerate the speed of chemical reactions without the expenditure of energy, is of the greatest importance in chemical industries, for the catalytic agent speeds up slow reactions, thus saving time and incidentally money—*e.g.* Deacon's process for chlorine; the sulphuric acid contact and chamber processes, etc. The contact process was the subject of a British patent by P. Phillips in 1831, but it was not until seventy years later, when R. Knietseh (1901) demon-

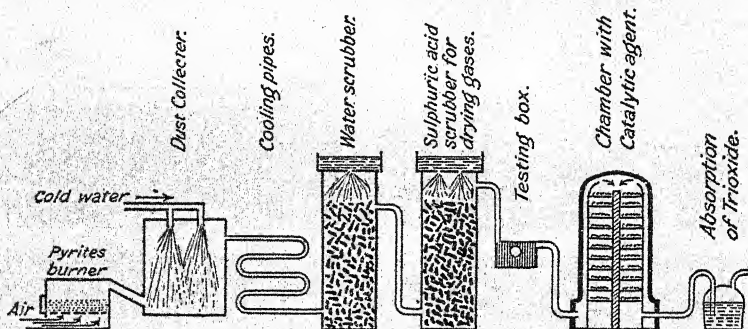


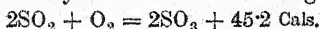
FIG. 178.—Diagrammatic Illustration of the Manufacture of Sulphuric Acid by the Contact Process.

strated the conditions necessary for successful work, that the contact process attained commercial success. The general disposition of the manufacturing plant is indicated in Fig. 178. The results are now so satisfactory that the "chamber process" is considered by many to be a "threatened industry" which will have to retire in favour of the more efficient "contact process." At first blush, the primary object of the industrial chemist is to make money; this may appear very sordid to the college-bred student; but it will be obvious that unless a manufacturing process realizes satisfactory profits, it must eventually collapse. Hence, the competition of rival processes frequently renders it necessary to overhaul old processes so as to make the methods more efficient and the management more economical. The chemistry of many industrial processes is comparatively simple when contrasted with the ingenuity displayed in handling large masses of material at a low cost, and in the conservation of by-products. Threatened industries sometimes live long, because competition of rivals acts as a healthy stimulus which leads to such improvements that the threatened industry has no difficulty in maintaining its position. Each rival incites the others to do their very best, and the world benefits.

1. The purification of the burner gases.—The mixture of sulphur dioxide and air from the "burners" must be very carefully cleaned from dust, finely divided sulphur, sulphuric acid, arsenic, etc., otherwise the platinized asbestos soon becomes inactive. Arsenic is particularly injurious. A very small trace of arsenic in the gases soon paralyzes—"poisons"—the activity of the platinum. The story of the contact process is essentially the story of the purification of the burner gases which have to be cooled, washed, and dried until the optical tests are satisfactory. The gases are cleaned by passage through a chamber—the "dust chamber"—in which steam is injected; through a series of lead pipes to reduce the temperature down to about 100°; through a series of towers in each of which the gases meet a descending spray of water. This washes the sulphuric acid, etc., from the gas. Finally, the gases are dried by passing them up a tower where they meet a descending stream of concentrated sulphuric acid. The gases are periodically tested for dust, arsenic, etc., to make sure the scrubbers, etc., are working efficiently.

2. The contact chamber.—The dried gases then pass into the contact chamber. In outline it consists of a cylinder which contains a number of tubes each packed with platinized asbestos placed between perforated shelves. The dry mixture of sulphur dioxide and air enters the contact chamber through a series of perforations near the base; and, after passing up to the top of the cylinder around the outside of the tubes containing the platinized asbestos, the gases descend through the platinized asbestos (where oxidation takes place), and leave the apparatus.

The platinized asbestos is arranged in this particular manner to avoid local rise of temperature by the heat evolved during the reaction:



Hence the system has a tendency to get hotter during the oxidation. Since the best results are obtained when the temperature of the asbestos is in the vicinity of 400°, everything is arranged so that the asbestos tubes are maintained at this temperature by the cold gases ascending in the interior of the cylinder.¹ Experience shows that 96 to 98 per cent. of the sulphur dioxide is oxidized in the cylinder.²

3. The condensation of the sulphur trioxide.—The "white mist" of sulphur trioxide is absorbed with great difficulty by water or dilute sulphuric acid; but it is rapidly and completely absorbed by 97 to 98 per cent. sulphuric acid. Hence the gases which leave the contact chamber pass into cast-iron tanks containing 97 to 98 per cent. sulphuric acid. A stream of water or dilute acid is run into the condensing tanks at such a rate that the strength of the acid is maintained at 97 to 98 per cent. If fuming acid be required, one or more wrought iron³ absorbing vessels are placed between the condensing tanks and the contact chamber.

§ 10. The Properties of Sulphuric Acid.

Contraction on dilution.—When sulphuric acid is mixed in water, the volume of the cold mixture is much less than the sum of the volumes of

¹ At the start, the asbestos is raised to the necessary temperature by heating the contact chamber from below.

² With ferric oxide as contact agent, 60 to 66 per cent. oxidation is obtained.

³ Experience shows that wrought iron resists the fuming acid more effectively than cast iron; and cast iron resists the more dilute acid better than wrought iron.

water and acid used. It is remarkable that the greatest contraction occurs with a solution containing about 97.7 per cent. of H_2SO_4 , that is, 79.7 per cent. of SO_3 . There is another maximum contraction with solutions containing 60 per cent. of free sulphur trioxide, that is a total per cent. of 92.65 SO_3 . A curve showing the volume of one gram of different solutions of sulphur trioxide and of sulphuric acid, H_2SO_4 , in water is shown in Fig. 179.

Boiling point.—Sulphuric acid (H_2SO_4) boils at about 270° with partial decomposition. Some sulphur trioxide passes off as vapour; the acid in the retort becomes weaker, and the boiling point steadily rises until the acid has attained a strength of about 98 per cent. H_2SO_4 (that is, 80 per cent. SO_3), when it distils over unchanged. Conversely, on boiling dilute solutions of sulphuric acid, the acid becomes stronger and stronger since water or very dilute acid passes over; at the same time, the boiling point steadily rises as illustrated in Fig. 180, until the acid has attained a strength of 98 per cent. H_2SO_4 , when it distils over unchanged, at 317° . The monohydrate fumes slightly, owing to the evolu-

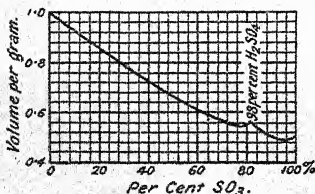


FIG. 179.—Contraction Curve.

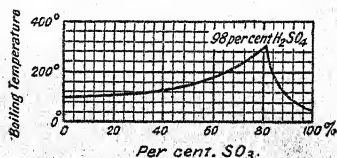
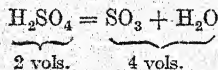


FIG. 180.—Boiling Curve.

tion of a little sulphur trioxide, when heated to above 30° . The facts just indicated are represented by the curve, Fig. 180, where the boiling-point curve shows a sharp apex at 317° . Below that temperature, water or dilute sulphuric acid, and above that temperature sulphur trioxide, is evolved. At that temperature (750 mm. pressure) a constant boiling acid is obtained. The specific gravity curve also shows a maximum, or the contraction curve a minimum (Fig. 179), with an acid of nearly that concentration. This acid also absorbs sulphur trioxide more rapidly—"greedily"—than either water or dilute sulphuric acid.

Vapour density.—The vapour of sulphuric acid is completely dissociated into water and sulphur trioxide at 450° . Granting this fact, and Avogadro's hypothesis, we have



The vapour density should be about half that required for H_2SO_4 . The theoretical value ($O = 32$) for H_2SO_4 is 98; H. St. C. Deville and L. Troost (1860) found 50.03, at 450° , which corresponds, very nearly, with a complete dissociation. At still higher temperatures, the sulphur trioxide itself dissociates, when heated, as indicated in the curve, Fig. 177. H. St. C. Deville and H. Debray (1860) suggested a process for the manufacture of oxygen on this reaction. They proposed to conduct the vapour of

sulphuric acid through a porcelain tube filled with bits of porcelain heated to bright redness. But nothing has come of the suggestion.

Freezing-point curves.—The freezing-point curves of solutions of sulphur trioxide in water are shown in Fig. 181. Maxima occur at points corresponding with $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$, $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$, H_2SO_4 , and $\text{H}_2\text{SO}_4 \cdot \text{SO}_3$ (pyrosulphuric acid). Our previous study of freezing-point curves (pp. 192 and 272) shows that these may be taken to represent the composition of definite compounds of sulphur trioxide and water formed under the conditions stated.

Hydrates of sulphur trioxide.—Several hydrates have been isolated: Pyrosulphuric acid— $\text{H}_2\text{O} \cdot 2\text{SO}_3$, or $\text{H}_2\text{SO}_4 \cdot \text{SO}_3$ —forms a transparent crystalline mass melting at 35° (p. 515). Monohydrated sulphuric acid— $\text{SO}_3 \cdot \text{H}_2\text{O}$, or H_2SO_4 —is a limpid colourless liquid without smell, with a specific gravity 1.850 at 15° , and 1.822 at 45° . It solidifies at 0° , forming crystalline plates which melt at 10° . Dihydrated sulphuric acid— $\text{SO}_3 \cdot 2\text{H}_2\text{O}$, or $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$ —crystallizes in hexagonal prisms, melting at 8° . The crystals are obtained by cooling 85 per cent. sulphuric acid with ice—particularly if the sulphuric acid has in solution

a large proportion of barium or lead sulphate. Pentahydrated sulphuric acid— $\text{SO}_3 \cdot 5\text{H}_2\text{O}$, or $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$ —melts at -25° . Trihydrated sulphuric acid— $\text{SO}_3 \cdot 3\text{H}_2\text{O}$, or $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$ —corresponds with a minimum point on the curve, and must therefore be regarded as a eutectic mixture under the conditions of the experiment (cf. p. 273). This does not mean that a hydrate $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$ will not exist under other conditions. Several claim to have isolated $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$ as a white crystalline mass, melting at about -70° .

The heat of dilution.—When sulphuric acid and water are mixed, a considerable rise of temperature occurs. The curve for the heat of solution rises steadily from 39 Cals. with 61.25 per cent. H_2SO_4 , to 193 Cals. with 100 per cent. H_2SO_4 ; and to 199 Cals. with sulphuric acid containing 2 per cent. of sulphur trioxide in solution, to 486 Cals. with 100 per cent. free sulphur trioxide. Hence sulphuric acid and water should be mixed with care. Do not pour the water into the acid, but always pour the acid into the water with constant stirring.

Absorption of water by sulphuric acid.—The great affinity of sulphuric acid for water is evidenced by its hygroscopicity and its use in desiccators, and in the balance case for maintaining a dry atmosphere in these vessels, and also for drying gases. Sulphuric acid acts upon solid and liquid substances depriving them of water; or even decomposing the substance—splitting off the elements of water when no ready formed water is present. Wood, paper, sugar, starch, and many organic substances are

¹ Many writers, probably more correctly, call $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$ the "monohydrate," etc.

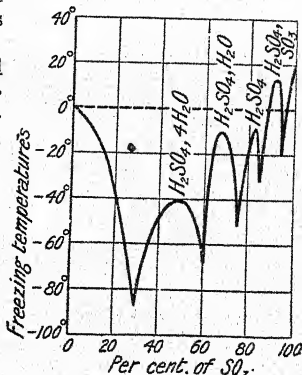


Fig. 181.—Freezing-point Curve.

blackened by concentrated sulphuric acid owing to the separation of carbon which accompanies the removal of the elements of water. This property is utilized (p. 834) for the preparation of carbon monoxide (p. 853); for the preparation of ethylene. The effect is easily demonstrated by stirring 10 grams of powdered cane sugar with 12 grams of concentrated sulphuric acid in a beaker. The sugar first becomes pale brown, rapidly darkens in tint, and finally becomes black; at the same time, much steam is evolved and the mass swells up considerably.

Action on metals.—Sulphuric acid is an oxidizing agent, particularly in hot concentrated solutions. Witness the preparation of sulphur dioxide by the action of hot concentrated sulphuric acid on copper, carbon, sulphur, etc. The cold concentrated acid does not perceptibly attack copper, mercury, antimony, tin, lead, or silver; but these metals are attacked by the hot concentrated acid. The prolonged action of sulphuric acid on mercury is said to give mercurous sulphide, mercuric sulphate, and sulphur dioxide; gold, platinum, iridium, and rhodium are not appreciably affected by the acid hot or cold. Dilute sulphuric acid dissolves magnesium, zinc, iron, cobalt, cadmium, manganese in the cold, forming a sulphate of the metal and hydrogen. The concentrated acid has very little action on these metals in the cold; a few bubbles of hydrogen may be evolved, but the action soon appears to stop. When heated, these metals give sulphur dioxide and the corresponding sulphates. Although the main reaction between zinc and dilute sulphuric acid can be represented by the equation $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$, it must be borne in mind that when sulphuric acid acts on zinc—particularly if the concentrated acid be used—the acid is partly reduced by the hydrogen (nascent), and sulphur dioxide, hydrogen sulphide, zinc sulphide, and free sulphur may be produced by the action of concentrated sulphuric acid on zinc. Warm, not cold, concentrated sulphuric acid is reduced by hydrogen, particularly in the presence of sulphates and the platinum metals. Impure hydrogen, even in the cold, reduces the acid, forming a trace of sulphur dioxide. It is necessary to bear this in mind when drying hydrogen with concentrated sulphuric acid.

Solubility of lead sulphate.—100 grams of sulphuric acid containing 95.6 per cent. of sulphuric acid dissolve 0.039 gram of lead sulphate; 100 grams of 86.0 per cent. sulphuric acid dissolve 0.011 gram; and 100 grams of 63.4 per cent. sulphuric acid dissolve 0.003 gram of lead sulphate. The amount of lead sulphate dissolved by more dilute solutions, say 2 per cent. sulphuric acid, can scarcely be detected. The diminishing solubility with increasing dilution is illustrated by the separation of lead sulphate as a white cloud when sulphuric acid, which has been in contact with lead vessels, is diluted with water.

§ 11. The Constitution of Sulphuric Acid and the Sulphates.

I am convinced that all theories on the constitution of organic compounds, and all controversies as to this or that mode of writing formulae, if not supported by a plausible arrangement of the compound atom, will aid us but little in the acquisition of correct ideas.—L. GMELIN (1848).

It will be obvious from our study of the constitution of sulphurous acid, that a chemical formula aims at representing the chemical nature of

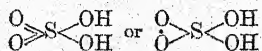
the compound in the simplest possible manner. The formula is intended to describe the way the compound behaves when it reacts with other compounds; and to show, at a glance, its relations with analogous compounds. Sulphuric acid, for example, is said "to contain two hydroxyl groups," meaning that in certain reactions, the OH groups can be exchanged for equivalent radicles; that there are certain relations between this acid and all other acids containing OH groups; etc. C. Schorlemmer has said that the structural formula shows "the past and future of a compound," that is, "the relation subsisting between its progenitors and its progeny." The formula for sulphuric acid is expressed $\text{SO}_2(\text{OH})_2$ for the following reasons:—

First.—Chlorine can react with concentrated sulphuric acid, forming chlorosulphonic acid $\text{Cl}-\text{HSO}_3$, where one OH group in sulphuric acid is replaced by chlorine. Phosphorus pentachloride PCl_5 —can displace two OH groups in sulphuric acid, forming sulphuryl chloride $\text{Cl}-\text{SO}_2-\text{Cl}$. Both these chloro-compounds react with water, forming sulphuric acid. The two OH groups can likewise be replaced with other radicles, *e.g.* NH_2 . Since two OH groups can be displaced together or separately, we infer from the rule of the constancy of structural arrangement that (1) sulphuric acid— H_2SO_4 —contains two hydroxyl—OH—groups.

Second.—Like sulphurous acid, it is possible to make but one compound of the type, $\text{CH}_3\text{O}-\text{SO}_2-\text{OK}$, by replacing the hydrogen of the hydroxyl groups with the radicles like CH_3 , C_2H_5 , etc. Hence it is inferred (2) the hydroxyl groups are related to the remainder of the atoms in the molecule H_2SO_4 in a symmetrical manner.

Third.—Certain univalent hydrocarbon radicles— C_2H_5 , C_6H_5 , etc.—can replace the chlorine in ClHSO_3 and in SO_2Cl_2 to form, say, ethyl sulphonic acid— $\text{C}_2\text{H}_5\cdot\text{SO}_2\cdot\text{OH}$, and diphenylsulphone— $(\text{C}_6\text{H}_5)_2\text{SO}_2$, respectively. The same compounds can be made by the oxidation of mercaptan— $\text{C}_2\text{H}_5\cdot\text{SH}$, and of diphenyl sulphide— $(\text{C}_6\text{H}_5)_2\text{S}$, in which the radicle must be joined directly to the sulphur atom. Assuming that the radicles remain fixed to the sulphur atom during the oxidation, it is inferred that (3) the hydroxyl groups in sulphuric acid— H_2SO_4 —are directly attached to the sulphur atom. Hence the formula $\text{HO}-\text{SO}_2-\text{OH}$ is preferable to $\text{HO}-\text{O}-\text{S}-\text{O}-\text{OH}$ for sulphuric acid.

Last.—We have not yet discussed whether the SO_2 group is constituted

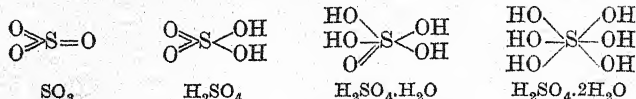


The sexivalency of sulphur in sulphur hexafluoride— SF_6 —points to the sexivalency of the sulphur atom in the sulphates, and hence it is probable that (4) each of the two oxygen atoms is attached to the sulphur atom by a double valency. For these reasons, the so-called rational formula of sulphuric acid is sometimes written $(\text{HO}_2)_2\text{SO}_2$, so that the constitutional formula is written:



always remembering that it is very probable that the best of our structural formulæ is not so closely related to the actual orientation of the atoms in the molecule as the stuffed and dried specimens of a museum are

related to the living organisms. The three hydrates of sulphur trioxide— H_2SO_4 , $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$, and $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$ —are sometimes written graphically:

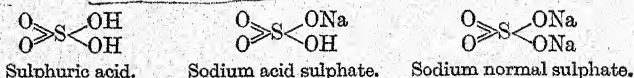


In harmony with the nomenclature employed for the periodic acids, the last formula represents the "maximum hydroxide" and corresponds with orthosulphuric acid, $\text{S}(\text{OH})_6$; the penultimate formula represents parasulphuric acid, $\text{SO}(\text{OH})_4$; and the ordinary formula for sulphuric acid therefore represents metasilphuric acid, $\text{SO}_2(\text{OH})_2$. There is not yet much experimental evidence in support of the two latter formulæ, and the method does not help us in dealing with the hydrate $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$; nor in understanding how these acids are all dibasic. Hence, it is doubtful if this is the correct interpretation of these hydrated forms of sulphuric acid.

Sometimes the following pedagogic distinctions are made between the different kinds of formulæ; (1) An empirical formula is the simplest mode of representing the composition of the substance as obtained by analysis—thus, HO represents the empirical formula of hydrogen peroxide; (2) the molecular formula represents the composition in harmony with the molecular weight—thus, H_2O_2 is the molecular formula of hydrogen peroxide; (3) a rational formula attempts to represent the chemical nature of the substance—thus, $(\text{HO})_2\text{SO}_2$ attempts to represent not only the composition molecularly, but it also indicates that it contains two replaceable hydroxyl groups; and (4) the structural, constitutional, graphic, or glyptic formula aims at representing the orientation of the atoms or groups of atoms in the molecule by a graphic structure.

§ 12. Sodium, Potassium, Mercury, and Copper Sulphates.

Sulphuric acid produces salts—sulphates—when it acts upon certain metals, metallic oxides, hydroxides, and carbonates. The sulphates are also prepared by heating salts of the more volatile acids—chlorides, nitrates, etc.—with sulphuric acid. Sulphuric acid is dibasic, forming two series of salts—normal and acid sulphates, p. 169:



The sulphates are usually fairly soluble in water, and crystallize readily. The sulphates of lead, calcium, strontium, and barium are but sparingly soluble in water.

Sodium and potassium sulphates.—The salts can be prepared by the general methods—action of acids on the corresponding bases, etc. While sodium sulphate crystallizes in monoclinic prisms with 10 molecules of water of crystallization, the potassium sulphate forms anhydrous rhombic prisms. The melting points of the anhydrous alkali sulphates are:

	Na_2SO_4	K_2SO_4	Rb_2SO_4	Cs_2SO_4
Melting points	888°	1070°	1051°	995°

The solubility of sodium sulphate has been discussed in connection with Fig. 58. Both sodium and potassium form acid sulphates—triclinic NaHSO_4 and rhombohedral KHSO_4 —when heated with sulphuric acid, and the

mixture allowed to cool. Acid sodium sulphate is less soluble in water than the corresponding potassium salt. Thus, 100 grams of water at 25° dissolve 55.3 grams of KHSO_4 , and 28.6 grams NaHSO_4 ; while at 100° the numbers are 121.6 and 50.0 respectively. Many minerals which resist attack by acids are broken down by fusion with the acid sulphates. When, say, acid potassium sulphate is fused, water is evolved and potassium pyrosulphate, $\text{K}_2\text{S}_2\text{O}_7$ —also written $\text{K}_2\text{SO}_4 \cdot \text{SO}_3$; or $\text{K}_2\text{O} \cdot 2\text{SO}_3$ —is formed: $2\text{KHSO}_4 = \text{H}_2\text{O} + \text{K}_2\text{S}_2\text{O}_7$. Some more complex acid salts have been reported: $\text{K}_2\text{O} \cdot 4\text{SO}_3$; and $\text{K}_2\text{O} \cdot 8\text{SO}_3$. Sodium does not form higher acid salts than the pyrosulphate. Sodium trihydrosulphate, NaHSO_4 , H_2SO_4 , has been indicated on p. 271; monohydrated trisodium hydrodisulphate, $\text{Na}_2\text{SO}_4 \cdot \text{NaHSO}_4 \cdot \text{H}_2\text{O}$, or $\text{Na}_3\text{H}(\text{SO}_4)_2 \cdot \text{H}_2\text{O}$, separates in rhombic prisms from a hot solution of an excess of sodium sulphate in sulphuric acid.

Manufacture of sodium and potassium sulphates.—(1) **Leblanc's salt-cake process.**—Large quantities of sodium sulphate are made in Leblanc's salt-cake process by warming sodium chloride with sulphuric acid in large cast-iron pans: $\text{NaCl} + \text{H}_2\text{SO}_4 = \text{NaHSO}_4 + \text{HCl}$. The torrents of hydrogen chloride which are given off are led through condensing towers—first devised by Gossage—and absorbed by water to form hydrochloric acid. After the first action is over, and the mixture in the pans begins to stiffen, the solid mass is raked on to the hearth of a reverberatory furnace and roasted by passing the hot gases from burning coke over the mass. In this way, most of the sodium acid sulphate is converted into the normal sulphate: $\text{NaHSO}_4 + \text{NaCl} = \text{Na}_2\text{SO}_4 + \text{HCl}$. The greater part of the hydrochloric acid from the sodium chloride comes from the pans—"pan gas"—and the remainder is evolved during the roasting—"roaster gas." The result of these operations is called "salt-cake." Salt-cake contains about 95 to 96 per cent. of normal sodium sulphate, and 4 per cent. of a mixture of sodium chloride, acid sodium sulphate, and other impurities which were associated with the original sodium chloride. Salt-cake is used in the manufacture of sodium carbonate. Potassium sulphate, if desired, can be made in the similar way from potassium chloride.

(2) **Hargreaves' process.**—This process for the conversion of sodium chloride into sodium sulphate avoids the intermediate formation of sulphuric acid. The gases from the pyrites burners—sulphur dioxide and air—are passed through a series of cast-iron cylinders in which the sodium chloride is heated to 500° or 560°. In a few days, the conversion of the chloride to sulphate is practically complete: $4\text{NaCl} + 2\text{SO}_2 + \text{O}_2 + 2\text{H}_2\text{O} = 2\text{Na}_2\text{SO}_4 + 4\text{HCl}$.

(3) **Sodium sulphate from kieserite.**—Sodium sulphate is made by mixing solutions of magnesium sulphate (from native kieserite) with sodium chloride. Double decomposition occurs: $2\text{NaCl} + \text{MgSO}_4 = \text{MgCl}_2 + \text{Na}_2\text{SO}_4$. On cooling the mixture, crystals of sodium sulphate are deposited from the solution.

(4) **Potassium sulphate from kainite.**—Potassium sulphate can be prepared from the crystals of $\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 6\text{H}_2\text{O}$ which remain when a hot saturated solution of kainite— $\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ —is cooled. The double sulphate of magnesium and potassium is treated with the necessary amount of potassium chloride; when crystals of normal

potassium sulphate separate from the solution; the mother liquid retains the more soluble double salt— KCl.MgCl_2 .

Mercurous and Mercuric sulphates.—These two salts can be made by heating mercury with sulphuric acid. Here, as is generally the case in preparing mercury salts, if the metal be in excess, mercurous salt is produced— Hg_2SO_4 ; and if the acid be in excess, the mercuric salt— HgSO_4 —is formed. Mercuric sulphate decomposes when heated, forming mercurous sulphate. Basic salts are formed when mercuric sulphate is brought in contact with water, *e.g.* the basic sulphate used in pharmacy under the name “turpeth mineral,” 2HgO.HgSO_4 , is formed as a yellow powder when mercuric sulphate is digested with boiling water.

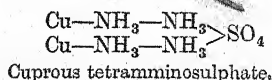
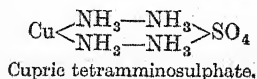
Cupric sulphate.—Cuprous sulphate— Cu_2SO_4 —has already been discussed, p. 464. When copper oxide or metallic copper is dissolved in sulphuric acid, blue triclinic crystals of copper sulphate— $\text{CuSO}_4.5\text{H}_2\text{O}$ —separate from the saturated solution on cooling. The crystals are also called “blue vitriol,” and they are prepared on a large scale by heating scrap copper in a furnace; sulphur is thrown on to the red-hot metal to form copper sulphide— Cu_2S ; air is then admitted, and the sulphide is oxidized to the sulphate. Copper sulphate is also made by roasting copper sulphide ores so that most of the iron is oxidized. The mass is then digested with “chamber sulphuric acid”; copper oxide dissolves much more quickly than iron oxide. The small amount of ferric sulphate found in the resulting solution is not objectionable for many of the applications of copper sulphate. Copper and iron sulphates cannot be separated by fractional crystallization because a double sulphate is deposited from solutions containing relatively large quantities of iron sulphate. To purify copper sulphate from iron, metallic copper is precipitated by dipping plates of clean iron into the solution; the resulting metal is dissolved in sulphuric acid, or calcined with sulphur as indicated above.

The properties of copper sulphate.—At 100° the blue crystals of $\text{CuSO}_4.5\text{H}_2\text{O}$ lose four molecules of water, and the residual monohydrate— $\text{CuSO}_4.\text{H}_2\text{O}$ —is bluish white. At 230° white anhydrous copper sulphate— CuSO_4 —is formed. The anhydrous sulphate is very hygroscopic, and it turns blue by the absorption of moisture. Hence anhydrous copper sulphate is used as a test for small amounts of water, and also for drying gases. More precise ideas on the dehydration of copper sulphate will be developed in the next section. Copper sulphate is used in electroplating, in electric batteries, dyeing, calico printing, preserving timber, as a germicide and insecticide for trees; and in the purification of certain types of drinking water.

Basic copper sulphates.—By digesting copper sulphate with copper oxide, for a long time in air, a yellow powder of $\text{CuSO}_4.\text{CuO}$ is formed. This basic sulphate, when thrown into cold water, forms a green insoluble compound— $\text{CuSO}_4.3\text{Cu(OH)}_2$; and in boiling water— $\text{CuSO}_4.2\text{Cu(OH)}_2$.

Ammonio- or ammino-copper sulphates.—When ammonia gas is passed into an aqueous solution of copper sulphate, the precipitate which first forms redissolves, and a deep blue solution is obtained when an excess of ammonia has been added. This solution deposits blue rhombic crystals, which, when dried over lime, have the composition $\text{CuSO}_4(\text{NH}_3)_4.\text{H}_2\text{O}$, and the name cupric tetramminosulphate. The crystals decompose on

exposure to the air, and, when heated to 149° , pass into $\text{CuSO}_4 \cdot 2\text{NH}_3$; at 203° , into $\text{CuSO}_4 \cdot \text{NH}_3$; and at 260° , into anhydrous CuSO_4 . The salt $\text{CuSO}_4 \cdot 5\text{NH}_3$ is formed when ammonia gas is brought in contact with anhydrous copper sulphate. The azure-blue solution produced when a solution of cupric sulphate is treated with an excess of aqueous ammonia contains cupric tetramminosulphate, and this is reduced to a colourless solution of cuprous tetramminosulphate, $\text{Cu}_2(\text{NH}_3)_4\text{SO}_4$, by hydrazine.



§ 13. The Vapour Pressure of Hydrates.

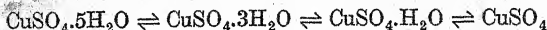
If a substance can form a number of definite hydrates, at any given temperature, each hydrate has its own characteristic vapour pressure. For instance, copper sulphate forms three hydrates with vapour pressures, at 50° .

Compound.	Formula.	Vapour Pressure.
Water alone	H_2O	92.0 mm.
Pentahydrate	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	47.0 mm.
Trihydrate	$\text{CuSO}_4 \cdot 3\text{H}_2\text{O}$	30.0 mm.
Monohydrate	$\text{CuSO}_4 \cdot \text{H}_2\text{O}$	4.5 mm.

Suppose that water vapour be slowly admitted to a "closed" vessel containing anhydrous copper sulphate at 50° , the vapour pressure will remain *constant and fixed* at 4.5 mm. until *all* the anhydrous copper sulphate has been transformed into the monohydrate: $\text{CuSO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{CuSO}_4 \cdot \text{H}_2\text{O}$; any further addition of water vapour will be followed by a rise in the vapour pressure which will remain constant and fixed at 30 mm. until all the monohydrate has been transformed into the trihydrate: $\text{CuSO}_4 \cdot \text{H}_2\text{O} + 2\text{H}_2\text{O} \rightleftharpoons \text{CuSO}_4 \cdot 3\text{H}_2\text{O}$. Further additions of water vapour will be attended by a rise of the vapour pressure to 47 mm. and the vapour pressure will remain stationary until all the trihydrate has been transformed into the pentahydrate: $\text{CuSO}_4 \cdot 3\text{H}_2\text{O} + 2\text{H}_2\text{O} \rightleftharpoons \text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. Any further addition of water vapour will raise the vapour pressure until the water condenses to a liquid and gives the vapour pressure of a saturated solution of copper sulphate at 50° . Further additions of water vapour will simply increase the amount of condensed water. When all the copper sulphate has dissolved, the vapour pressure will be determined by the concentration of the solution of copper sulphate.

Conversely, starting with the pentahydrate, the reverse change will occur if water be continually withdrawn from the system. The vapour pressure of the pentahydrate, 47 mm. at 50° , will be maintained as long as the system contains any of the pentahydrate. When all the pentahydrate has decomposed, the vapour pressure will drop suddenly to 30 mm. and remain stationary until all the trihydrate has decomposed; the vapour pressure will then fall abruptly to 4.5 mm. and remain at that value

until all the monohydrate has decomposed into anhydrous copper sulphate. These changes are usually shown diagrammatically by curves resembling Fig. 182. Amounts of water, expressed in gram-molecules, are represented on the horizontal axes, and the vapour pressures along the vertical axes. The constancy of the vapour pressure of each hydrate is emphasized by the horizontal terraces on the vapour pressure curve. In symbols, omitting water molecules :



This step-by-step dissociation of the hydrates furnishes a method which is sometimes available for deciding whether or not definite compounds exist at definite temperatures. If definite compounds are produced, the gradual addition or removal of water vapour will alter the vapour pressure curve until a pressure is reached which remains constant for a certain period, and then suddenly assumes a new constant value. It must be added that some (*e.g.* P. Blackman, 1911) consider the dehydration of copper sulphate pentahydrate is comparable with the removal of water from an ordinary aqueous solution, that the "breaks" in the curve are due to a mal-interpretation of imperfect experiments, and that the vapour

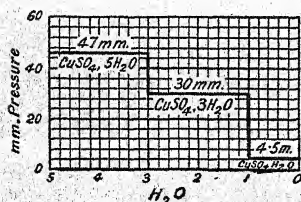


FIG. 182.—Vapour Pressure Curves of the Hydrates of Copper Sulphate.

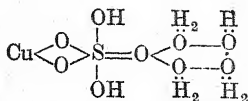
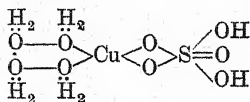
compounds of ammonia with silver chloride, copper sulphate, copper chloride, etc.

pressures in the preceding table are points on a continuous curve. This, however, does not interfere with the principle involved. By measuring the rate of decomposition of hydrated aluminium and ferric hydroxides, W. Ramsay inferred the non-existence of definite hydrates; but he concluded that two were formed with lead oxide— $2\text{PbO} \cdot \text{H}_2\text{O}$, and $3\text{PbO} \cdot \text{H}_2\text{O}$. J. M. van Bemmelen, for similar reasons, believed in the non-existence of definite hydrates of silicic acid. The principles just described can also be applied to

§ 14. Water of Crystallization.

The study of copper sulphate, which is typical of numerous other hydrates, shows that although the molecules of the combined water may differ in the tenacity with which they are retained by the molecule of copper sulphate, yet the water of one hydrate does not differ in kind from that of the other hydrates. This water is sometimes conventionally styled "constitutional water," "water of crystallization," "water of hydration," or "water of combination." The actual term used does not matter very much provided it is not misunderstood. The mode of writing the formulæ— $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$; $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ —and the ease with which the hydrates dissociate into water, etc., might give rise to the idea that the water molecule exists in the hydrate ready-made. There is, however, no evidence how the elements of water are combined in the hydrate, and consequently, perhaps following the old idea, "truth is within ourselves, and takes no rise from outward things," suggestions have been made to represent

hydrates graphically, thus, pentahydrated cupric sulphate have been represented by



and other forms. In both formulæ just indicated, it is assumed (i) that one water molecule is associated in the molecule differently from the other four molecules because one molecule requires a higher temperature for its expulsion than do the other four molecules; and (ii) that two of the other four molecules are associated differently from the other two because of the step-by-step dissociation. We really know very little beyond the simple facts that: (1) Water is a product of the dissociation of the hydrates; (2) water is usually given off at comparatively low temperatures; (3) water is not an essential part of the reacting unit in its most characteristic transformations; and (4) water is not generally necessary for the formation of the salt since water of crystallization can generally be removed by suitable means leaving the salt anhydrous. Several zeobites may lose their combined water without losing their crystalline form.

If the term "water of crystallization" be carelessly employed, it may suggest that crystallization is somehow dependent on the presence of water, and this the more, as efflorescent salts "appear" to lose their crystalline character when water is lost. Crystals of gypsum— $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ —form a white chalky powder when the water is driven off; crystalline sodium carbonate, and also Glauber's salts, likewise produce white powders when their combined water is expelled. The powdered dehydrated substances are all crystalline. In fact, practically all chemical compounds can be crystallized. Crystallization is not dependent upon the presence of water. Sulphur, common salt, iodine, potassium chlorate, potassium sulphate, and numerous other crystalline substances do not contain the elements of water. Again, crystalline calc spar does not contain the elements of water, and yet when calcined it gives a white powder. The calc spar has lost carbon dioxide, not water.

To prevent any misunderstanding, the student should here make sure that he is perfectly clear about the use of the terms anhydrous and anhydride. The latter term, in inorganic chemistry, is applied to those oxides which unite with water to form an acid; while "anhydrous" is an adjective applied to oxides, salts, etc., when it is desired to emphasize the fact that they do not contain the elements of water—water of crystallization, water of combination, etc.

Alcohol, $\text{C}_2\text{H}_5\text{OH}$, has a constitution similar to water, but one of the hydrogen atoms of water is replaced by the radicle C_2H_5 , and in ether both are replaced:



Water.



Ethyl alcohol.



Ethyl ether.

Both alcohol and hydrogen peroxide can combine with certain other molecules to form complexes, and thus we speak of "alcohol of crystallization," and "hydrogen peroxide of crystallization." Similarly, in view of the ammonio sulphates of copper, we might speak of ammonia of

crystallizations; thus cobalt chloride forms $\text{CoCl}_2 \cdot 6\text{NH}_3$ and $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$; and $\text{CoCl}_2 \cdot 2\text{NH}_3$ and $\text{CoCl}_2 \cdot 2\text{H}_2\text{O}$.

§ 15. Calcium, Barium, Strontium, Magnesium, Zinc, Ferrous, and Lead Sulphates

Calcium sulphate.—This salt is a by-product in some chemical processes. It occurs in nature in white masses as *gypsum*— $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$. It is used as a fertilizer under the name “land plaster,” and also in the manufacture of plaster of Paris. If the gypsum occurs in clean fine-grained masses, it is called *alabaster*; and if in colourless translucent crystals, *selenite*. An anhydrous form of calcium sulphate also occurs native in rhombic crystals, it is called *anhydrite*— CaSO_4 . There is also a variety of anhydrous calcium sulphate which is made artificially, and which is more soluble than native anhydrite. Artificial anhydrite is an unstable variety.

A. S. Marggraf (1746) found that when potassium or calcium sulphate is heated with charcoal, a gas with a pungent smell characteristic of burning sulphur is given off, and hence he inferred that these substances are compounds of sulphuric acid; and in 1765 A. L. Lavoisier proved that gypsum is a compound of sulphuric acid and lime, and later analyses of the purest varieties show that gypsum is composed of 79.1 per cent. of calcium sulphate, and 20.9 per cent. of water. When these numbers are transferred into a molecular ratio, by dividing the weight by the molecular weights of calcium sulphate and water respectively, the empirical formula for gypsum corresponds with that indicated above— $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$.

The action of water on calcium sulphate.—The solubility of calcium sulphate varies in an unusual manner with changes of temperature, for it gradually increases up to 38° , and then diminishes as illustrated in Fig. 183.

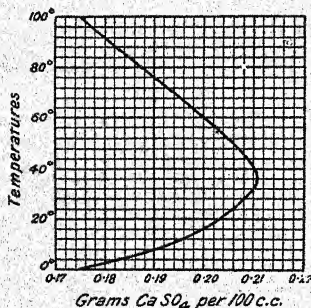


Fig. 183.—Solubility of Calcium Sulphate.

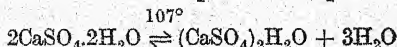
A hydrate in the presence of its solution is stable if it be less soluble than the other hydrates (or the anhydrous salt); in other words, if a salt forms several hydrates, that hydrate will be stable which is in equilibrium with the weakest solution. For instance, below 32.8° , a solution saturated with $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ is less concentrated than a solution saturated with the anhydrous salt, and accordingly, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ is the stable compound; above 32.8° the solution of anhydrous sodium sulphate is less concentrated than a solution of $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$, and therefore the anhydrous salt is the more stable form.

Again, a hydrate will be stable in the presence of a solution if the vapour pressure of the hydrate be less than the vapour pressure of the solutions with which it is in equilibrium. If the vapour pressure be greater, the solid phase changes to the hydrate possessing the next lowest vapour pressure. The hydrate with the largest proportion of water and the smallest

vapour pressure will be formed in the coldest solution. For instance, above 32.8° , the vapour pressure of $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ in contact with anhydrous sodium sulphate is greater than the vapour pressure of the solution saturated with anhydrous sodium sulphate, and therefore Glauber's salt, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$, changes into the solid anhydrous salt and its saturated solution. The vapour pressure of an aqueous solution of anhydrite, CaSO_4 , at ordinary temperatures is greater than the vapour pressure of an aqueous solution of gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$. Hence, gypsum will be deposited from such a solution, and solid anhydrite placed in the solution will pass into gypsum.

The usual effect of a foreign salt on a solution is to lower the vapour pressure of the solution at a given temperature. This may cause the formation of a lower hydrate. Thus the vapour pressure of solutions of calcium sulphate containing much sodium and magnesium chlorides is lower than the vapour pressure of a solution of gypsum, and accordingly, anhydrite is deposited from such solutions; again, if gypsum be placed in the solution, it will pass into anhydrite. For instance, gypsum in contact with pure water changes into anhydrite at about 66° , but in contact with a saturated solution of sodium chloride, the inversion temperature, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O} = 2\text{H}_2\text{O} + \text{CaSO}_4$, is as low as 30° ; while if other salts are present, the inversion temperature may be still lower. This explains the formation of anhydrite when the Stassfurt deposits were laid by the natural evaporation of sea-water. If anhydrite does occur in nature, it generally occurs in deep-seated regions protected from moist air. If it occurs near the surface, and is so brought in contact with moisture, it is always associated with gypsum into which it is slowly changing. The presence of anhydrite in the Stassfurt deposits was inexplicable until these facts were discovered. The absence of the hemihydrate was also puzzling until it was found that the hemihydrate is in an unstable condition if other more stable forms are present, when it slowly changes into gypsum or anhydrite if one of these salts be in contact with it. The hemihydrate is stable under a wide range of conditions if the more stable anhydrite or gypsum be absent.

Plaster of Paris.—When gypsum is heated to about 120° it loses the equivalent of $1\frac{1}{2}$ molecules of water, and forms a hemihydrate, $(\text{CaSO}_4)_2\text{H}_2\text{O}$, as a white powder called *plaster of Paris*—because of the large deposits of gypsum employed for the manufacture of plaster at Montmartre (Paris). The transformation or inversion temperature is represented:



If gypsum be heated with water under pressure at 150° , "silky" needle-like crystals of the hemihydrate are formed. Hence this hydrate is said to occur in "boiler deposits" when water containing gypsum in solution is heated in steam boilers. The hemihydrate is unstable as indicated above. Plaster of Paris is a mixture of the anhydrous, dihydrate, and hemihydrate, and hence in virtue of the presence of the hemihydrate, it too is an unstable system. The change of the hemihydrate into anhydrite and gypsum is so very slow at ordinary temperatures that the setting qualities are not lost by prolonged storage out of contact with air.

When plaster of Paris is wetted with, say, one-third of its weight of water, it forms a plastic mass which "sets" in from 5 to 15 minutes to

a white, porous, hard mass. Plaster of Paris is a valuable material for making exact reproductions of statues, ornaments, moulds, decorative work, etc. The slight expansion which occurs during the setting of plaster of Paris enables it to make a sharp reproduction of the details of a mould. Different grades are used for wall plasters, cements, paper sizings, etc. Admixtures of borax, alum, etc., with the plaster retard the rate of setting; while common salt, etc., accelerate the rate of setting. Alum makes "set" plaster much harder—the mixture of plaster with alum is called *Keene's cement*; while a mixture with borax is called *parian cement*; and a mixture of plaster with hydraulic lime is called *Scott's selenitic cement*.

Plaster of Paris, in setting, reforms a mass of interlacing needle-like crystals of the dihydrate. Good plaster of Paris probably contains a mixture of the dihydrate, hemihydrate, and the anhydrous salt. Anhydrous calcium sulphate is more soluble than the dihydrate, and accordingly, when water is mixed with the plaster, the anhydrite dissolves first, and passes into the dihydrate. This recrystallizes about the undissolved particles of the dihydrate in needle-like crystals.

If plaster of Paris be heated above 200° , it passes into anhydrous calcium sulphate which does not harden and set because it takes up water very slowly. This is the so-called "dead-burnt" plaster. A. Potylitzin called this β - CaSO_4 to distinguish it from the α - CaSO_4 , formed between 120° to 200° , which readily takes up water and hardens. According to Rohland, another soluble form, similar to, if not the same as α - CaSO_4 , is formed when gypsum is heated to about 520° . This is the "Estrich-Gyps" or "flooring plaster" of the Germans, so named because it is largely used in Germany for making floors. The slower rate of setting of the hard-fired plasters is probably due to (1) delayed recrystallization of the dihydrate owing to supersaturation; and (2) to the slower rate of solution of the anhydrite calcined at a high temperature. The texture of the "set" plaster is thus dependent on the rate of setting, which is in turn dependent upon the composition of the plaster, that is, on the temperature of calcination, impurities in the native gypsum used in making the plaster, etc.

A series of double sulphates of calcium and the alkali metals M have been prepared—syngenites, $\text{CaM}_2(\text{SO}_4)_2 \cdot \text{H}_2\text{O}$. Potassium syngenite, $\text{CaK}_2(\text{SO}_4)_2 \cdot \text{H}_2\text{O}$, occurs in Stassfurt salts. Ammonium syngenite is formed in silky needles when calcium sulphate is added to a nearly saturated solution of ammonium sulphate.

Barium and strontium sulphates.—Native strontium sulphate— SrSO_4 —is called *celestine*. Barium sulphate occurs in nature as *barytes* or *heavy spar* which occurs in large rhombic prisms. Crystals of anhydrite and of the barium and strontium sulphates are deposited from solutions in hot concentrated sulphuric acid. Barium, strontium, and lead (not calcium) sulphates are isomorphous. They form mixed crystals with one another, but calcium sulphate does not form mixed crystals with the others. Changes analogous with those observed with the hydrates of calcium sulphate have not been observed with barium and strontium sulphates. Calcium and barium sulphates are reduced by heating in a stream of dry carbon monoxide between 900° and 950° ; the reduction commences at about 680° or 700° . Reduction with carbon in an atmosphere of nitrogen commences at about 700° , and is completed at 1000° ;

at low temperatures, the carbon is converted into carbon dioxide, and at high temperatures into carbon monoxide. Barium and strontium sulphates are formed as heavy white precipitates when sulphuric acid, or a soluble sulphate, are added to solutions of barium or strontium salts respectively. The precipitates are very sparingly soluble in water. Comparing the solubilities of barium, strontium, and calcium sulphates: 100 grams of water at 10° dissolve 0.19 gram of CaSO_4 —calcium sulphate; 0.10 gram of SrSO_4 —strontium sulphate; and 0.0002 gram BaSO_4 —barium sulphate. The low solubility of barium sulphate is utilized for the detection and determination of both barium and sulphuric acid. If a soluble sulphate be added to a soluble barium salt, the barium will be precipitated as barium sulphate; and conversely, if a soluble barium salt be added to a soluble sulphate, the "sulphate" will be precipitated as barium sulphate. Barium sulphate is fairly soluble in hot concentrated sulphuric acid, and the solution, on cooling, deposits an acid barium sulphate— $\text{BaSO}_4 \cdot \text{H}_2\text{SO}_4$. The solution also deposits the normal sulphate as a white precipitate when diluted with water. Double sulphates of the type $\text{SrSO}_4 \cdot \text{K}_2\text{SO}_4$, and $\text{SrSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4$, have been reported.

Barium sulphate is used as a source of the barium salts—one method of transformation was discussed in connection with barium chloride; in a second process, the barium sulphate is fused with sodium carbonate: $\text{BaSO}_4 + \text{Na}_2\text{CO}_3 = \text{BaCO}_3 + \text{Na}_2\text{SO}_4$. The sodium sulphate can be leached out with water, and barium carbonate remains as a residue. This can be dissolved in the acid necessary for the production of the required salt. If the product of the reaction be treated with an acid, before washing out the sodium sulphate, the barium salt, formed by the action of the acid on the carbonate, will at once react with the soluble sulphate and regenerate insoluble barium sulphate. Barium sulphate is employed under the name "permanent white" in the manufacture of paint, as a filling for writing paper, etc.

Magnesium sulphate, MgSO_4 .—Magnesium sulphate occurs as *kieserite*, $\text{MgSO}_4 \cdot \text{H}_2\text{O}$, in the Stassfurt deposits. When *kieserite* is digested with water, and the solution purified by recrystallization, colourless rhombic prisms of the heptahydrate— $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ —separate from the cold solution. When the temperature is about 30°, monoclinic prisms of the hexahydrate— $\text{MgSO}_4 \cdot 6\text{H}_2\text{O}$ —are formed. Several hydrates are revealed when magnesium sulphate is studied by the method employed for ferric chloride. At 150° the crystalline salt passes into the monohydrate, *kieserite*, $\text{MgSO}_4 \cdot \text{H}_2\text{O}$; and at 200° the anhydrous salt is obtained with partial decomposition. Magnesium sulphate forms a series of double salts— $\text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$, *schönite*—crystallizing in monoclinic prisms, isomorphous with $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$. Like barium sulphate, crystals of the acid salt $\text{MgSO}_4 \cdot \text{H}_2\text{SO}_4$ are deposited from a hot solution of magnesium sulphate in sulphuric acid; and from cold solutions, the acid salt $\text{MgSO}_4 \cdot 3\text{H}_2\text{SO}_4$ is crystallized. Crystalline magnesium sulphate— $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ —is also called "Epsom salts," or *epsomite*, because it occurs abundantly in the water of the Epsom springs. This salt is used in medicine; in the tanning and dyeing industries; and in the manufacture of paints and soaps.

Ferrous sulphate, FeSO_4 .—The rhombic prisms of $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ are isomorphous with the corresponding salts of zinc— $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$; $\text{BeSO}_4 \cdot 7\text{H}_2\text{O}$; cadmium— $\text{CdSO}_4 \cdot 7\text{H}_2\text{O}$; iron— $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$; manganese—

$\text{MnSO}_4 \cdot 7\text{H}_2\text{O}$; chromium— $\text{CrSO}_4 \cdot 7\text{H}_2\text{O}$; cobalt— $\text{CoSO}_4 \cdot 7\text{H}_2\text{O}$; and nickel— $\text{NiSO}_4 \cdot 7\text{H}_2\text{O}$. It is interesting to note that copper sulphate normally crystallizes, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, but a salt isomorphous with the above series, $\text{CuSO}_4 \cdot 7\text{H}_2\text{O}$, is formed when copper sulphate in admixture with the sulphate of iron, zinc, or magnesium, is allowed to crystallize. Ferrous sulphate— $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ —is also called "green vitriol" and "copperas." The term "copperas" appears to be a corruption of the French "*couperose*," which, in turn, is a corruption of the Latin "*cupri-rosa*"—literally "rose of Cyprus." Cyprus was once renowned for its copper mines. The German word for copperas is "vitriol," a term used by Pliny, about 50 A.D. Pliny described this substance as "vitriolus quasi vitrum," meaning "vitriol, a kind of glass," since the crystals resembled green glass, but differed from glass in being easily soluble in water. Sulphuric acid was obtained from "vitriol," and Geber accordingly gave it an Arabic name, which, translated into English, means oil of vitriol. The term "oil" was applied because the acid flows sluggishly like olive oil, and it has an oily feel when rubbed between the fingers (this latter experiment is dangerous). Ferrous sulphate is used in the manufacture of blue pigments; as a mordant; and in the manufacture of black inks.

Ferrous sulphate can be made by dissolving iron in sulphuric acid, and also by exposing iron pyrites (marcasite) to air and moisture. Oxidation occurs, and the liquid which drains away contains ferrous and ferric sulphates and sulphuric acid; the "drainage solution" is converted into ferrous sulphate by the addition of scrap iron; on crystallization, the solution furnishes pale green crystals of ferrous sulphate— $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$. Crystals with 1, 2, 3, 4, 5, and 6 molecules of water of crystallization have been reported. Aqueous solutions of the salt oxidize quickly when exposed to the air; the crystals also gradually turn brown, presumably owing to the formation of a basic ferric sulphate, say, $\text{Fe}_2\text{O}(\text{SO}_4)_2$, which does not dissolve completely in water. The double salt— $\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$ —ferrous ammonium sulphate, or, Mohr's salt, is employed in volumetric analysis. Its aqueous solution, particularly if acidified with sulphuric acid, oxidizes much more slowly than ferrous sulphate.

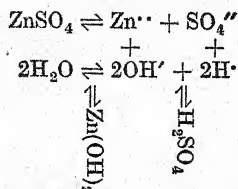
Zinc sulphate, $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$.—This salt is very like the magnesium sulphate in its general behaviour. It is made by roasting the native sulphide in air. The zinc sulphate which is formed at the same time is extracted with water—*vide* p. 37. The salt is called "white vitriol." Zinc sulphate is used in making certain battery solutions.

The hydrolysis of zinc and magnesium salts.—Reference might now be made to the hydrolysis of zinc and magnesium chlorides, p. 287. The electrical conductivity of the purest water yet made is very slight, but it can be detected; it is 0.000000038, this number means that a column of water one centimetre long will not conduct an electrical current so well as a column of mercury of equal cross section and extending more than 30 times round the earth. Accordingly, it is inferred that the purest water yet made is slightly ionized: $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}'$, and that water is thus an amphoteric oxide. Eighteen grams of such water contains about $\frac{1}{1000000}$ of a gram of H^+ ions, and $\frac{17}{10000000}$ of a gram of OH' ions. If all or part of one or other of these two ions be removed, more water molecules will ionize in order to keep the equilibrium value constant.

The ionization of a few weak acids deduced from conductivity measurements for $\frac{1}{10}$ N-solutions, with HCl as standard of reference, is

Hydrochloric acid.	91.400
Carbonic acid	0.174
Hydrosulphuric acid	0.075
Boric acid	0.013
Hydrocyanic acid.	0.011

Zinc sulphate is completely ionized in dilute solution, $\text{ZnSO}_4 \rightleftharpoons \text{Zn}^{++} + \text{SO}_4^{--}$. The Zn^{++} ions pair with the OH' ions of the water to form feebly ionized $\text{Zn}(\text{OH})_2$ molecules; more water is ionized in order to maintain the equilibrium value, $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}'$; and these actions continue until the concentration of the zinc hydroxide has attained the equilibrium value, $\text{Zn}(\text{OH})_2 \rightleftharpoons \text{Zn}^{++} + 2\text{OH}'$. When that occurs, a considerable proportion of the OH' molecules have been withdrawn from the solution to form molecules of zinc hydroxide, and an excess of the H^+ ions in solution are "paired," so to speak, with the SO_4^{--} ions of the zinc sulphate. The scheme may be represented:



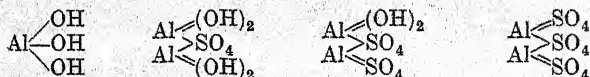
An aqueous solution of zinc sulphate thus behaves in many ways like a solution of sulphuric acid—*e.g.* turns blue litmus red, etc. If an alkali, say, sodium hydroxide, be added to the solution of zinc sulphate, the OH' ions of the base unite directly with the H^+ ions of the acid to form water, and if just sufficient OH' ions be added to remove all the H^+ ions of the acid, the solution will contain nothing more than would be obtained by dissolving sodium sulphate in water; the zinc hydroxide is removed from the solution by precipitation when its concentration exceeds the solubility coefficient.

Lead sulphate, PbSO_4 .—The mineral *anglesite* is the native form of lead sulphate. Its rhombic crystals are isomorphous with strontium and barium (not calcium) sulphates. It is formed as a white precipitate when sulphuric acid or a soluble sulphate is added to a solution of a lead salt. *Vide p. 526.* A solution of lead sulphate in sulphuric acid, when boiled, deposits acid lead sulphate— $\text{PbSO}_4 \cdot \text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$. Lead sulphate dissolves in hot concentrated hydrochloric acid, in ammonium acetate, and in sodium thiosulphate. When potassium sulphate is added to a soluble lead salt, a double lead potassium sulphate, $\text{PbK}_2(\text{SO}_4)_2$, is formed. To prevent the precipitation of this salt when lead is precipitated as sulphate in analysis, the precipitation is made in boiling solutions. An ammonium lead sulphate has also been prepared. The basic lead sulphate $2\text{PbSO}_4 \cdot \text{PbO}$ is made commercially by reducing galena with carbon, and subliming the product in a stream of air. It is used as a white pigment—*sublimed white lead*—and is not so readily discoloured as white lead.

§ 16. Aluminium Sulphate and the Alums.

Aluminium sulphate, $\text{Al}_2(\text{SO}_4)_3$.—This salt is prepared by dissolving the hydrated oxide in sulphuric acid. Large quantities too are made by dissolving bauxite, and the purer varieties of clay in the same acid. Clay roasted at a dull red heat is more readily attacked by the acid. The crude aluminium sulphate so obtained is called "alum-cake," and if much iron is present, "alum ferric cake," used in the purification of sewage. A purer sulphate is made by heating bauxite with sodium carbonate, or by boiling cryolite with milk of lime. In each case a solution of sodium aluminate— $\text{Al}_2\text{O}_3 \cdot 3\text{Na}_2\text{O}$ —almost free from iron is obtained. The sodium aluminate is then decomposed by a current of carbon dioxide, and the precipitated aluminium hydroxide is dissolved in sulphuric acid. When the solution is concentrated, the mass solidifies to a white solid which does not crystallize very readily. The crystalline sulphate has the composition represented by: $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$.

The aqueous solution of aluminium sulphate has an acid reaction, and in the presence of zinc it evolves hydrogen; a basic sulphate is formed at the same time. Obviously, the aluminium sulphate is hydrolyzed in aqueous solution: $\text{Al}_2(\text{SO}_4)_3 + 6\text{H}_2\text{O} \rightleftharpoons 2\text{Al}(\text{OH})_3 + 3\text{H}_2\text{SO}_4$. The basic sulphates are formed by boiling the sulphate with freshly precipitated aluminium hydroxide. The composition of the aluminium sulphates—aluminium trivalent—can be represented graphically:



Al-hydroxide. Dibasic sulphate. Monobasic sulphate. Normal sulphate.

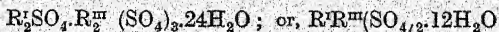
One bivalent oxygen atom can be introduced in place of two univalent hydroxyl groups, when the anhydrous basic sulphates are in question as illustrated by the basic ferric sulphates below. The mineral *keramohalite* or *alunogen* is a monobasic sulphate, $\text{Al}_2\text{O}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$, found in Chili. Werner has another mode of viewing the constitution of the alums which will be described later.

Alums.—When a hot solution of aluminium sulphate is mixed with potassium sulphate, and the solution is cooled, octahedral crystals of a



FIG. 184.—Abnormally developed Alum Crystals.

double sulphate of aluminium and potassium separate. The octahedral form is not always recognizable directly on account of the abnormal development of one or more faces. Thus, shapes like Fig. 184 may be obtained. Cubic crystals of alum can be obtained by adding a little ammonia or alkali carbonate to a solution of ordinary alum at 30° to 40° , and allowing it to cool. If the liquid be heated to a higher temperature, a basic sulphate may be deposited, and ordinary crystals of alum formed as the solution cools. The salt has the empirical formula, $\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$, or else $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$. This salt is a typical member of a large number of isomorphous compounds which are called "alums." Their general formula is:



where R^I represents an atom of a univalent metal or radicle—potassium, sodium, ammonium, rubidium, caesium, silver, and univalent thallium; and R^{III} represents an atom of a trivalent metal—aluminium, iron, chromium, manganese, cobalt, rhodium, indium, gallium, and trivalent thallium. An iridium rubidium alum has also been reported. The report of a silver alum has not been confirmed. Soda alum is formed in well-defined octahedral crystals by mixing supersaturated solutions of the two component sulphates at ordinary temperatures and then cooling the liquid; or by mixing neutral solutions of the component sulphates, evaporating to a specific gravity 1.38 to 1.42, and cooling the liquid. If the crystallization starts above 30° , small crystals of aluminium sulphate are formed, but these give way to the crystals of soda-alum on standing some weeks. An alum is a compound which crystallizes with 12 molecules of water, and is derived from two molecules of sulphuric acid by replacing one hydrogen atom by a univalent atom or radicle; and the remaining three hydrogen atoms by a trivalent metal:

Potash alum	$KAl(SO_4)_2 \cdot 12H_2O$
Ammonia alum	$NH_4Al(SO_4)_2 \cdot 12H_2O$
Iron alum	$KFe(SO_4)_2 \cdot 12H_2O$
Ammonium ferric alum	$NH_4Fe(SO_4)_2 \cdot 12H_2O$
Rhodium thallium alum	$TlRh(SO_4)_2 \cdot 12H_2O$
Caesium manganese alum	$CsMn(SO_4)_2 \cdot 12H_2O$
Chromium potash alum	$KCr(SO_4)_2 \cdot 12H_2O$

Potash-alum or kalinite occurs native as the result of the weathering of iron pyrites and alkali rocks. Mixed crystals of different alums are readily obtained by allowing mixed solutions of different alums to crystallize; and crystals of one alum, say chrome alum, grown on the surface of ordinary alum, or conversely, can be obtained. A series of selenium alums have also been made in which the sulphur atom in K_2SO_4 is replaced by an atom of selenium. The selenium alums are isomorphous with the alums proper.

The stability of the alums generally decreases with increasing atomic weight of the trivalent metal, and increases in stability with increasing atomic weight of the univalent metal. For example, sodium forms alums only with the lightest trivalent metals—aluminium, vanadium, and chromium; potassium forms stable alums only with aluminium and chromium. The ferric and gallic alums readily form basic salts; and indium does not form a potash-alum, but gives instead the double sulphate $K_2SO_4 \cdot In_2(SO_4)_3 \cdot 6$ (and 8) H_2O ; titanium forms an alum only with caesium and rubidium. Lithium does not form a well-defined "alum," and soda alum is not easy to make. The alums are usually much more soluble in hot water than in cold; and they are deposited from cooling concentrated solutions of the two salts in the form of octahedral or cubical crystals. The solubilities of a number of alums have been compared by J. Locke, and his results are, at ordinary temperatures:

	K	NH_4	Ti	Rb	Cs
Aluminium	0.28	0.39	0.18	0.06	0.01 mols per litre.
Vanadium	—	1.21	0.57	0.18	0.02 " "
Chromium	0.44	0.41	0.21	0.08	0.02 " "
Ferric (iron)	—	1.66	0.80	0.29	0.05 " "
Indium	—	—	—	—	0.17 " "

With a given trivalent metal, the higher the atomic weight of the alkali

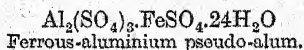
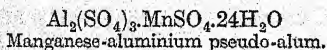
metal the less the solubility of the corresponding alum. Ammonia alum is rather more soluble than potash-alum. Thallium alums come between potash- and rubidia-alums. In passing through the tervalent elements, the solubility of the alum formed with a given alkali metal increases with the atomic weight of the tervalent element. The difference is more marked with the lighter univalent metals. The chrome alums coming between those of vanadium and iron show the same general behaviour as the other alums, but they do not fall in with the solubility rule for increasing atomic weights. The molecular volumes and melting points of such alums as do not decompose also fall into the same relative order as their solubilities.

The solubilities and melting points of the alums of the alkali metals are compared in the following scheme :

	Sodium.	Potassium.	Ammonium.	Rubidium.	Cæsium.
Melting point	66°	92.5°	94.5°	105°	120°
Solubility 16°, 100 grams water	51	15	12	2.2	0.6

When heated, the alums give up their water of crystallization. Potash-alum melts at about 92.5°, and slowly loses all its water of crystallization at 100°; at higher temperatures, the water is driven off more rapidly, and leaves behind a white porous mass called "burnt alum." Ammonia alum when calcined gives a residue of alumina— Al_2O_3 —which may contain traces of sulphates. The soluble sulphates can be removed by washing.

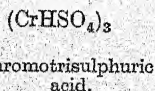
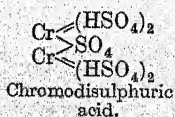
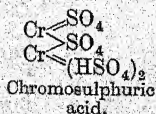
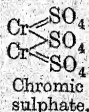
Pseudo-alums.—A series of double sulphates, sometimes called pseudo-alums, can be made by introducing a bivalent element—manganese, ferrous iron, copper, zinc, magnesium—in place of the univalent element of ordinary alums. Thus,



The latter is related to the mineral *halotrichite*. These alums are not isomorphous with ordinary alums.

Alum-stone or alunite is a kind of basic alum— $\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 4\text{Al}(\text{OH})_3$ —found near Rome, in Hungary, etc. It is supposed to have been formed by the action of volcanic sulphur dioxide on the felspathic rocks. It is insoluble in water, but, on calcination, it gives a residuum of alumina (mixed with ferric oxide as impurity) and potash-alum passes into solution when the mass is digested with water. Alum prepared in this way is called "Roman alum," and that which occurs in commerce is crystallized in cubes. Roman alum is fairly free from iron, and it has been in demand—in dyeing, etc.—where a specially pure alum is required; but the modern methods of preparation furnish an alum quite as free from iron.

Chromic sulphate, $\text{Cr}_2(\text{SO}_4)_3$, forms alums isomorphous with those produced by ferric and aluminium sulphates. With sulphuric acid, chromic sulphate forms a series of acid sulphates—chromosulphuric acids :



A chromopyrosulphuric acid with the radicle HS_2O_7 , has been prepared, $\text{Cr}(\text{OH})(\text{HS}_2\text{O}_7)_2$.

Ferric sulphate, $\text{Fe}_2(\text{SO}_4)_3$.—This salt is prepared by adding sulphuric acid and an oxidizing agent—nitric acid, hydrogen peroxide, etc.—to ferrous sulphate. Thus, $2\text{FeSO}_4 + \text{H}_2\text{SO}_4 + \text{H}_2\text{O}_2 = \text{Fe}_2(\text{SO}_4)_3 + 2\text{H}_2\text{O}$. If the solution be boiled, a basic sulphate is precipitated. On evaporation at a low temperature, a white mass of anhydrous salt, $\text{Fe}_2(\text{SO}_4)_3$, mixed with a little basic sulphate is formed. If the solution of ferric sulphate be mixed with the necessary amount of potassium sulphate, violet octahedral crystals of iron alum, mentioned above, separate. The formulæ of some basic ferric sulphates are:



Ferric sulphate.



Ferric oxysulphate.



Ferric dioxysulphate.



Ferric oxide.

A series of hydroxybasic sulphates are represented graphically by formulæ similar to those employed for the basic aluminium sulphates.

Cobaltic sulphate, $\text{Co}_2(\text{SO}_4)_3$, when a solution of cobaltous sulphate and sulphuric acid is electrolyzed in the anode compartment of a divided cell, with dilute sulphuric acid about the cathode, crystals of cobaltic sulphate are formed. When the salt is treated with potassium or ammonium sulphate, blue octahedral crystals, isomorphous with the alums, are formed. Both the cobalt alums are unstable, readily decomposing in aqueous solution with the evolution of oxygen.

§ 17. Persulphuric Acid and the Persulphates.

M. Berthelot (1878) found that when a mixture of oxygen and sulphur dioxide is exposed to the silent discharge in an ozone tube, oily drops of sulphur heptoxide, S_2O_7 , are formed: $2\text{SO}_2 + \text{O}_3 = \text{S}_2\text{O}_7$. The liquid forms white crystals at 0° . It gradually decomposes, on keeping, into sulphur trioxide and oxygen. Sulphur heptoxide combines with water with a hissing noise like sulphur trioxide, but the solution has not the same properties as if sulphur trioxide alone had been dissolved in the water. A similar solution can be obtained by mixing concentrated sulphuric acid with hydrogen peroxide in the cold, and by the electrolysis of concentrated sulphuric acid.

The preparation of persulphates.—It will be remembered that when dilute sulphuric acid is electrolyzed, hydrogen and oxygen are obtained in the proportion: 2 volumes of hydrogen, and one volume of oxygen. As the concentration of the acid is increased, less and less oxygen is evolved; until, with 50 per cent. sulphuric acid and a cold solution, inappreciable quantities of oxygen will be disengaged at the anode. A new compound is formed—persulphuric acid, $\text{H}_2\text{S}_2\text{O}_8$. The best way of studying this action is to electrolyze a saturated solution of potassium sulphate in sulphuric acid, specific gravity about 1.3.

The solution is placed in a test-tube, *A*, Fig. 185, so that the test-tube is about three-fourths filled. The test-tube is fixed in a large beaker by means of the wire *C*. A glass cylinder, *D*, is fixed in the test-tube by means of the wire *B*. A coil of platinum wire—the cathode—is allowed to dip into the solution of potassium sulphate as indicated in the diagram; and a platinum wire sealed to a

piece of glass tube so that about $1\frac{1}{2}$ to 2 cm. of the wire projects from the tube forms the anode. This is fixed within the cylinder. The large beaker is filled with water in which pieces of ice are floating—"iced-water." A current of about one or two amperes is sent through the solution. Hydrogen appears at the cathode; and a white crystalline mass accumulates at the anode. In about 45 minutes the current may be stopped, the white crystals of potassium persulphate collect on a filter paper; wash with alcohol; and finally with ether.

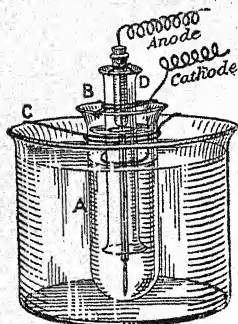


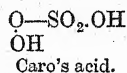
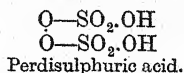
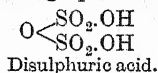
FIG. 185.—Preparation of Potassium Persulphate.

The formula $K_2S_2O_8$ is conformable with the change in the molecular conductivity of solutions of the salt with dilution, and with the lowering of the freezing point of solutions of the salt. H. Marshall, who first isolated potassium persulphate in 1891, suggested the formula KSO_4 , but later favoured the doubled formula given above.

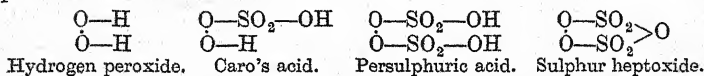
The properties of persulphates.—The solid persulphates are fairly stable. Potassium persulphate decomposes on heating into potassium pyrosulphate and oxygen: $2K_2S_2O_8 = 2K_2S_2O_7 + O_2$. An aqueous solution of a persulphate is a powerful oxidizing agent, and is used for that purpose in analytical work, e.g. it precipitates a brown manganese peroxide from soluble manganese salts. The persulphates *slowly* liberate iodine from solutions of potassium iodide; oxidize iodine to iodic acid, etc. A solution of a manganese salt gives a precipitate of manganese peroxide with potassium persulphate; but, in the presence of silver nitrate, is oxidized to pink permanganate—*Marshall's reaction*. The intensity of the coloration is proportional to the amount of manganese present, and hence the reaction is utilized for the colorimetric determination of manganese. The silver salt is said to act catalytically presumably by the continuous formation and reduction of silver peroxide. Unlike hydrogen peroxide, the persulphates do not give orange-yellow colorations with titanium salts. Most of the persulphates are soluble. Barium persulphate is readily soluble in water. Hence barium chloride gives no precipitate with the persulphates, although it does give a precipitate with a sulphate, but if the persulphate be decomposed by warming, a precipitate of barium sulphate is obtained. A dilute solution of the acid—persulphuric acid, also called perdisulphuric acid—can be made by treating the barium persulphate with sulphuric acid. The persulphates are used in photography for "reducing" negatives; and ammonium persulphate is used in technical organic chemistry.

Caro's acid.—If potassium persulphate be digested with 40 per cent. sulphuric acid in a freezing mixture so that there is no rise of temperature; or if concentrated sulphuric acid and hydrogen peroxide (5 per cent.) be mixed together, a permonosulphuric acid is obtained: $H_2SO_4 + H_2O_2 \rightleftharpoons H_2O + H_2SO_5$. The solution is sometimes called Caro's acid, H_2SO_5 —after its discoverer, N. Caro, 1898. An acid, 92.3 per cent. purity, has been made by the action of pure hydrogen peroxide on sulphur trioxide: $H_2O_2 + SO_3 = H_2SO_5$, or on sulphuric acid free from water. The equilibrium condition $H_2O_2 + H_2SO_4 \rightleftharpoons H_2SO_5 + H_2O$ is reached when about two-thirds of the theoretical monopersulphuric acid is formed. The pure acid forms a white crystalline mass which melts at about 45° , and it

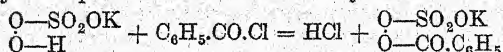
is comparatively stable. Like persulphuric acid, Caro's acid has strong oxidizing qualities. Unlike persulphuric acid, it liberates iodine from potassium iodide *at once*. It also oxidizes sulphur dioxide to the trioxide; ferrous salts to ferric salts; hydrogen chloride to chlorine, but it is without action on hydrogen fluoride; and precipitates peroxides from salts of silver, copper, manganese, cobalt, and nickel. It does not bleach permanganates, nor oxidize chromic nor titanium salts, and is therefore distinct from hydrogen peroxide. It decomposes explosively in contact with finely powdered platinum or silver, but not with lead or zinc dust. Cellulose and wool are instantly decomposed by the acid. All the metals examined, except gold and platinum, react with persulphate solutions, either passing directly into solution, or remaining undissolved in the form of oxides or basic salts. No gas is evolved in the case of zinc, nickel, etc. The relations of the two acids with pyro- or di-sulphuric acid, will appear from the supposed graphic formula:



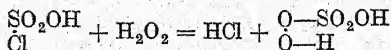
Persulphuric acid is accordingly sometimes called perdisulphuric acid to distinguish it from "permonosulphuric acid" or Caro's acid. Another scheme brings out the relation of these persulphuric acids to hydrogen peroxide:



There has been some discussion whether Caro's acid is monobasic, H_2SO_5 , or dibasic, $\text{H}_2\text{S}_2\text{O}_9$. The analysis of the potassium salt is not conclusive, since KHSO_5 would have the same ultimate composition as the salt $\text{K}_2\text{S}_2\text{O}_9 \cdot \text{H}_2\text{O}$. Benzoyl chloride, $\text{C}_6\text{H}_5\text{COCl}$, reacts with the potassium salt of Caro's acid, forming the benzoyl derivative. This reaction undoubtedly corresponds with the monobasicity of the acid:

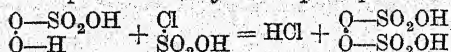


It is very unlikely that this result would occur if the formula of the salt in question were $\text{K}_2\text{S}_2\text{O}_9 \cdot \text{H}_2\text{O}$. The synthesis of Caro's acid by the action of the calculated quantity of chlorosulphonic acid, $\text{SO}_2\text{Cl}(\text{OH})$, on pure hydrogen peroxide in the cold:



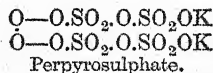
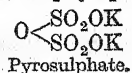
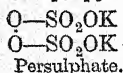
is in agreement with the view of the molecular formula just indicated.

Persulphuric acid.—If Caro's acid be mixed with the calculated quantity of chlorosulphonic acid, crystals of persulphuric acid are formed:



The hydrogen chloride is removed by keeping the mixture in a desiccator under reduced pressure. The white crystalline mass so obtained is more stable than Caro's acid. It melts just above 60° with decomposition. The aqueous solution is first hydrolyzed into Caro's acid and sulphuric acid: $\text{H}_2\text{S}_2\text{O}_8 + \text{H}_2\text{O} = \text{H}_2\text{SO}_5 + \text{H}_2\text{SO}_4$; and then decomposes into

oxygen and sulphuric acid: $2\text{H}_2\text{S}_2\text{O}_8 + 2\text{H}_2\text{O} = 4\text{H}_2\text{SO}_4 + \text{O}_2$. In consequence, the pure acid cannot be prepared by treating the persulphates with dilute acids, as indicated above, and then concentrating the solution. Hydrogen peroxide is made technically either by distilling the solution of persulphuric acid obtained by the electrolysis of sulphuric acid, or by distilling potassium persulphate with dilute sulphuric acid. When sodium or potassium persulphate is exposed to the action of sulphur trioxide vapour two molecules of the latter unite with one molecule of the former, producing potassium perpyrosulphate, $\text{KS}_2\text{O}_8 \cdot \text{O}_2\text{S}_2\text{O}_6\text{K}$.

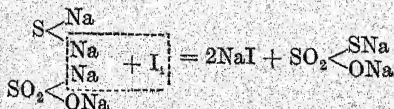


The perpyrosulphates deliquesce in air without the evolution of oxygen, and the solution contains both persulphuric acid and hydrogen peroxide; in direct contact with water the salt dissolves with a hissing noise giving off ozonized oxygen.

§ 18. Thiosulphuric Acid and the Thiosulphates.

Preparation.—If an aqueous solution of sodium sulphite, Na_2SO_3 , be exposed to the air, one oxygen atom per molecule of sodium sulphite is taken up, and sodium sulphate, Na_2SO_4 , is formed: $2\text{Na}_2\text{SO}_3 + \text{O}_2 = 2\text{Na}_2\text{SO}_4$. Similarly, if sodium sulphite be digested with finely divided sulphur for some time, one atom of sulphur per molecule of sodium sulphite is taken up, and a new salt, sodium thiosulphate, $\text{Na}_2\text{S}_2\text{O}_3$, is formed: $\text{Na}_2\text{SO}_3 + \text{S} = \text{Na}_2\text{S}_2\text{O}_3$. These reactions suggest some analogy in the structure of the thiosulphates and the sulphates; and this is emphasized by the term *thio*-sulphates, meaning *sulpho*-sulphates—from the Greek *θειον* (theion), sulphur. The term *sodium hyposulphite* formerly applied to sodium thiosulphates obviously meant “a salt derived from a sulphur acid containing less oxygen than sulphurous acid, or less oxygen per molecule than a sulphite; with the sodium salt, such a substance would have the formula Na_2SO_2 .” So long as such a salt was unknown, no confusion could arise, but when Schützenberger’s acid was discovered, § 4, it became necessary to apply a term other than hyposulphite to the photographer’s “hyposulphite” in order to keep the notation consistent.

Sodium thiosulphate is also formed when a mixture of sodium sulphide and sulphite is treated with iodine: $\text{Na}_2\text{S} + \text{Na}_2\text{SO}_3 + \text{I}_2 = \text{Na}_2\text{S}_2\text{O}_3 + 2\text{NaI}$. This reaction is sometimes supposed to occur in two steps: $\text{Na}_2\text{S} + \text{I}_2 = 2\text{NaI} + \text{S}$; and the liberated sulphur acts upon the sodium sulphite as indicated above. The net result of the reaction is that the iodine withdraws one atom of sodium from the molecule of sodium sulphide and one from the sodium sulphite, while the residues unite to form a more complex molecule condensation product. This operation is sometimes called Spring’s reaction—after W. Spring’s syntheses of the thionic acids by this reaction in 1874:



The thiosulphates are also formed by the action of sodium sulphide— Na_2S or Na_2S_2 —upon sulphur dioxide. Some sulphur separates at the same time: $2\text{Na}_2\text{S} + 3\text{SO}_2 = 2\text{Na}_2\text{S}_2\text{O}_3 + \text{S}$. The action is supposed to occur in three steps: (1) formation of sodium sulphite and hydrogen sulphide: $\text{SO}_2 + \text{H}_2\text{O} + \text{Na}_2\text{S} = \text{Na}_2\text{SO}_3 + \text{H}_2\text{S}$; (2) the hydrogen sulphide reacts with the sulphur dioxide, as indicated on p. 497, liberating free sulphur: $\text{SO}_2 + 2\text{H}_2\text{S} = 2\text{H}_2\text{O} + 3\text{S}$; and (3) the liberated sulphur is taken up by the sodium sulphite as indicated above. By boiling sulphur with milk of lime or with sodium hydroxide, a mixture of the corresponding thiosulphate and sulphide is formed, *e.g.* $3\text{Ca}(\text{OH})_2 + 12\text{S} = \text{CaS}_2\text{O}_3 + 2\text{CaS}_5 + 3\text{H}_2\text{O}$. The calcium sulphide is converted into the thiosulphate on oxidation by exposure to the air: $2\text{CaS}_5 + 3\text{O}_2 = 6\text{S} + 2\text{CaS}_2\text{O}_3$. When the "tank waste" of Leblanc's process (*q.v.*) is exposed to the air, the calcium sulphide, CaS , is in part transformed into calcium thiosulphate, CaS_2O_3 , the latter, when treated with sodium carbonate forms insoluble calcium carbonate, and soluble sodium thiosulphate: $\text{CaS}_2\text{O}_3 + \text{Na}_2\text{CO}_3 = \text{Na}_2\text{S}_2\text{O}_3 + \text{CaCO}_3$.

Thiosulphuric acid.—Thiosulphuric acid has not been isolated. By acting on a thiosulphate with a mineral acid, thiosulphuric acid appears to be formed, but it begins to decompose at once into sulphur dioxide and free sulphur: $\text{Na}_2\text{S}_2\text{O}_3 + 2\text{HCl} = 2\text{NaCl} + \text{H}_2\text{O} + \text{SO}_2 + \text{S}$. The sulphur only appears after the lapse of a certain time—seconds or minutes, according to the concentration of the solution. It has been stated that the decomposition of the thiosulphate does not occur at once. If, however, the acidified solution be neutralized before the turbidity appears, the neutralization does not stop the separation of the sulphur. Some thiosulphuric acid must therefore have decomposed. Probably the very finely divided sulphur is not visible until the fine particles have clotted into larger granules. The evolution of sulphur dioxide with the separation of sulphur on the addition of a dilute mineral acid distinguishes thiosulphates from sulphites in qualitative analysis.

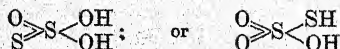
Properties and uses of thiosulphates.—Sodium thiosulphate crystallizes in large transparent monoclinic prisms with five molecules of water of crystallization: $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$, the phenomenon of undercooling exhibited by this salt, has been previously discussed. Sodium thiosulphate is the most important salt of thiosulphuric acid, and it is called, by photographers, sodium hyposulphite, or simply "hypo"; they have used it extensively for fixing ever since J. Herschel discovered the solvent action of the alkaline salts on silver chloride. This salt is readily soluble in water, and the aqueous solution readily dissolves silver mercurous, or thallous, chloride, bromide, or iodide. A soluble silver sodium thiosulphate is formed: $2\text{AgCl} + 3\text{Na}_2\text{S}_2\text{O}_3 \rightleftharpoons \text{NaCl} + 2\text{Ag}_2\text{Na}_4(\text{S}_2\text{O}_3)_2$. Hence its use in the "fixing bath" of the photographer. Solutions of the thiosulphates give a precipitate of silver sulphide with silver nitrate; presumably silver thiosulphate is first formed and immediately decomposed; $\text{Ag}_2\text{S}_2\text{O}_3 + \text{H}_2\text{O} = \text{H}_2\text{SO}_4 + \text{Ag}_2\text{S}$; in contrast, sulphites with silver nitrate give a precipitate of silver sulphite which is decomposed only when boiled. Hence the amount of thiosulphate in a mixture of the two salts can be determined by adding silver nitrate and subsequently titrating with $\frac{1}{20}$ N-alkali, using methyl orange as indicator. Crystallized sodium thiosulphate decomposes when heated, forming hydrogen sulphide, sulphur, and

sodium sulphite and sulphate. Sodium stannite, $\text{Sn}(\text{ONa})_2$, reacts with sodium thiosulphate, forming, not sulphate, but rather black stannous sulphide: $2\text{Sn}(\text{ONa})_2 + \text{Na}_2\text{S}_2\text{O}_3 + \text{H}_2\text{O} = \text{SnS} + \text{Na}_2\text{SO}_3 + \text{Na}_2\text{SnO}_3 + 2\text{NaOH}$, corresponding with the decomposition of thiosulphuric acid: $2\text{H}_2\text{S}_2\text{O}_3 = 2\text{SO}_2 + 2\text{H}_2\text{S} + \text{O}_2$.

Sodium thiosulphate is readily oxidized by potassium permanganate, nitric acid, chlorine, etc. Thus, with ferric chloride, the violet colour first produced is slowly discharged, and the ferric salt is reduced to the ferrous state: $\text{Na}_2\text{S}_2\text{O}_3 + 8\text{FeCl}_3 + 5\text{H}_2\text{O} = 2\text{NaHSO}_4 + 8\text{FeCl}_2 + 8\text{HCl}$. Under certain conditions, the thiosulphates can yield oxygen to powerful reducing agents, and they thus become oxidizing agents, *e.g.* sodium thiosulphate is reduced by sodium amalgam reforming sodium sulphide and sulphite: $\text{Na}_2\text{S}_2\text{O}_3 + 2\text{Na} = \text{Na}_2\text{SO}_3 + \text{Na}_2\text{S}$. Sodium thiosulphate is often used in preference to sulphurous acid (*q.v.*) as an "antichlor" in order to remove the last trace of chlorine from the bleached goods. The action depends upon the reducing qualities of sodium thiosulphate: $\text{Na}_2\text{S}_2\text{O}_3 + 5\text{H}_2\text{O} + 4\text{Cl}_2 = \text{Na}_2\text{SO}_4 + \text{H}_2\text{SO}_4 + 8\text{HCl}$; or $\text{Na}_2\text{S}_2\text{O}_3 + \text{H}_2\text{O} + \text{Cl}_2 = \text{Na}_2\text{SO}_4 + 2\text{HCl} + \text{S}$. If the thiosulphate be used as an antichlor for photographic printing paper, the last traces of the antichlor are exceedingly difficult to wash away, and the trace which remains may cause a fading of the print. Hence, certain "hypo eliminators" have been recommended. These are usually oxidizing agents whose products are not so deleterious as the hypo itself. Thus Newton (1855) recommended alum; F. W. Hart (1866), sodium hypochlorite; Smith and Spiller (1866), hydrogen peroxide; Vogel (1872), iodine; and Schering (1895), potassium persulphate, sold under the trade name *anthion*.

Solutions of iodine are quickly decolorized by sodium thiosulphate with the formation of sodium tetrathionate, $\text{Na}_2\text{S}_4\text{O}_6$. Thus, $2\text{Na}_2\text{S}_2\text{O}_3 + \text{I}_2 = \text{Na}_2\text{S}_4\text{O}_6 + 2\text{NaI}$. If a solution of sodium thiosulphate of known strength be added from a burette to a solution containing some iodine, until the colour of the iodine has just disappeared, the amount of thiosulphate required for the work of decolorization furnishes the datum necessary for calculating the amount of iodine in the given solution.

Constitution of thiosulphates.—It is not very clear whether thiosulphuric acid, and accordingly also the thiosulphates, should be represented by graphic formula based on:

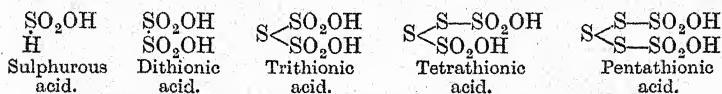


Some prefer the latter, others the former; and some consider that both varieties exist. The former is then called symmetrical, and the latter asymmetrical thiosulphuric acid. The available evidence, however, favours the asymmetrical $\text{HS.SO}_2\text{OH}$. It is known that phosphorus sulphide acts on many compounds so as to substitute sulphur for oxygen, and generally, the oxygen of the hydroxyl groups are more susceptible to attack than oxygen united to a sulphur atom. Since thiosulphuric acid is formed by the action of phosphorus sulphide on sulphuric acid, it is inferred that the product of the action will contain a HS-group. Again, the formation of thiosulphuric acid by the action of hydrogen sulphide on sulphur trioxide: $\text{H}_2\text{S} + \text{SO}_3 = \text{HO.SO}_2\text{SH}$, is considered to be analogous to the formation of sulphuric acid by the action of water on

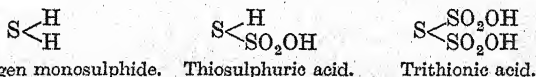
sulphur trioxide: $\text{H}_2\text{O} + \text{SO}_3 = \text{HO.SO}_2.\text{OH}$. Hence the formula $\text{HS.SO}_2.\text{OH}$ is usually preferred for thiosulphuric acid.

§ 19. Polythionic Acids.

There is a remarkable series of five acids—called collectively the polythionic acids—closely related to sulphurous acid and to thiosulphuric acid. The polythionic acids include di-, tri-, tetra-, penta-, and hexathionic acids. To show C. W. Blomstrand's (1869) and D. I. Mendeléeff's (1870) views of the inter-relations of the polythionic acids it is convenient to consider the group SO_2OH —i.e. $\text{O}=\text{S}(\text{OH})_2$ as a monad radicle. The constitution of the thionic acids is then represented by the following graphic formulæ—with hydrogen $\text{H}-\text{H}$ as the starting-point:

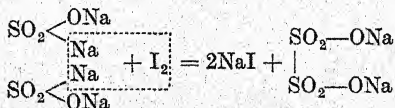


Dithionic acid is therefore a disulphonic acid with asymmetric sulphurous acid as the corresponding monosulphonic acid. We previously encountered this curious faculty of sulphur, whereby "chains of atoms" can be linked together, in our study of the hydrogen sulphides. From this point of view, pentathionic acid is related to hydrogen trisulphide; tetrathionic acid to hydrogen persulphide; and trithionic acid and thiosulphuric acid are related to hydrogen monosulphide as follows:



The two latter are respectively mono- and di-sulphonic acids of hydrogen sulphide. It is interesting to compare the latter with pyrosulphuric acid, persulphuric acid, and Caro's acid previously discussed, since in these compounds, oxygen takes the place of sulphur. There are other modes of interpreting the known properties of these acids, but Blomstrand's views, just indicated, agree best with more recent observations.

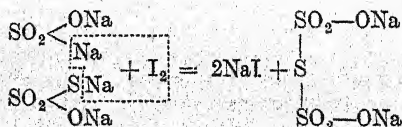
Dithionic acid, $\text{H}_2\text{S}_2\text{O}_6$.—The sodium salt is made together with some sulphuric acid, by Spring's reaction with iodine on sodium sulphite:



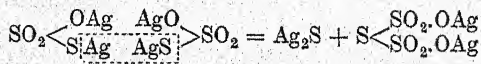
Dithionic acid has been called "hyposulphuric acid," and the salts, "hyposulphates." The manganese salt, MnS_2O_6 , is made by passing sulphur dioxide through water with manganese dioxide in suspension: $\text{MnO}_2 + 2\text{SO}_2 = \text{MnS}_2\text{O}_6$. At the same time there is a side reaction between the manganese dithionate and the manganese dioxide: $\text{MnO}_2 + \text{MnS}_2\text{O}_6 = 2\text{MnSO}_4$. This reduces the yield. Ferric hydroxide or cobaltic hydroxide forms ferrous or cobaltous sulphate and dithionate when treated with sulphurous acid at a low temperature. Manganese dithionate is converted into

barium dithionate, BaS_2O_6 , by mixing it with barium hydroxide, $\text{Ba}(\text{OH})_2$ and the resulting salt can be purified by crystallization: $\text{BaS}_2\text{O}_6 \cdot 2\text{H}_2\text{O}$ is formed. This gives soluble dithionic acid and insoluble barium sulphate when treated with dilute sulphuric acid. The aqueous solution of the acid can be concentrated by evaporation until its specific gravity is about 1.35, any further concentration decomposes the acid: $\text{H}_2\text{S}_2\text{O}_6 = \text{SO}_2 + \text{H}_2\text{SO}_4$. If the sodium dithionate be reduced in aqueous solution by means of sodium amalgam, sodium sulphite is reformed: $\text{Na}_2\text{S}_2\text{O}_6 + 2\text{Na} = 2\text{Na}_2\text{SO}_3$. Some consider the constitution of dithionic acid to be $\text{H}.\text{SO}_3.\text{O}.\text{SO}_2.\text{H}$ related with pyrosulphuric acid, $\text{HO}.\text{SO}_2.\text{O}.\text{SO}_2.\text{OH}$. Some dithionic acid as well as sulphuric acid is said to be formed when sulphur dioxide is passed into a solution of potassium permanganate or dichromate. Ozone oxidizes both thiosulphates and sulphites into dithionates and sulphates. Lead, barium, magnesium, and sodium peroxides do not yield dithionates when treated with sulphurous acid.

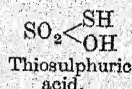
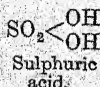
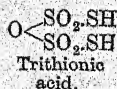
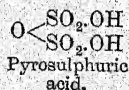
Trithionic acid, $\text{H}_2\text{S}_3\text{O}_6$.—The potassium salt of this acid is formed by passing sulphur dioxide through a concentrated solution of potassium thiosulphate: $3\text{SO}_2 + 2\text{K}_2\text{S}_2\text{O}_3 = \text{S} + 2\text{K}_2\text{S}_3\text{O}_6$. There is some doubt if Spring's reaction, the action of iodine upon a mixed solution of sodium sulphite and thiosulphate, proceeds:



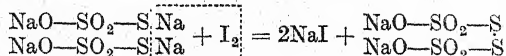
as might be expected; because a mixture of the sulphate and tetrathionate is obtained. Sodium trithionate is also formed by the action of sulphur chloride, SCl_2 , upon sodium sulphite; and by warming an aqueous solution of potassium acid sulphite with flowers of sulphur: $6\text{KHSO}_3 + 2\text{S} = 2\text{K}_2\text{S}_3\text{O}_6 + \text{K}_2\text{S}_2\text{O}_3 + 3\text{H}_2\text{O}$. By boiling silver thiosulphate with water, a molecule of Ag_2S splits from two molecules of the thiosulphate and silver trithionate remains:



The acid itself is formed from the potassium salt by the addition of hydrofluosilicic acid. Potassium fluosilicate is precipitated, and the trithionic acid remains in solution. The acid and its salts are readily decomposed into sulphur, and sulphuric acid or a sulphate. By the reducing action of sodium amalgam, sodium trithionate is converted back into sodium sulphite and sodium thiosulphate. According to Chancel and Diacon, when a solution of sodium dithionate is boiled with sodium sulphide, sodium thiosulphate is formed. There is thus an intimate relation between thiosulphuric and trithionic acids. Some try to emphasize this relationship by stating that trithionic acid is related to pyrosulphuric acid with water, the way that thiosulphuric acid is related to sulphuric acid:

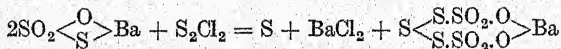


Tetrathionic acid, $\text{H}_2\text{S}_4\text{O}_6$.—The sodium salt is formed by Spring's reaction with iodine and sodium thiosulphate:



The barium salt is prepared in a similar manner, and from this the acid itself is obtained by the action of dilute sulphuric acid. The reaction under consideration is the last of the set of condensations by the removal of an atom of sodium from each of two molecules of a salt and the condensation of the residues to form a more complex molecule—with sodium thiosulphate, di-, (tri-), and tetrathionate. All these reactions are reversed and the original salts reproduced by treating the complex salts with sodium amalgam. Tetrathionic acid can be made by passing a current of hydrogen sulphide into sulphurous acid until the smell of sulphur dioxide is discharged. The solution can be warmed on a water-bath to get rid of sulphur dioxide and hydrogen sulphide; the water lost by evaporation being replaced from time to time. The solution of thiosulphuric acid is filtered to get rid of sulphur. A 15 per cent. solution of an alkali or alkaline earth hydroxide reacts with a tetrathionate forming a sulphite and thiosulphate: $2\text{Na}_2\text{S}_4\text{O}_6 + 6\text{NaOH} = 3\text{Na}_2\text{S}_2\text{O}_3 + 2\text{Na}_2\text{SO}_3 + 3\text{H}_2\text{O}$; with a concentrated solution (1:1) of the hydroxide, some sulphide is also formed: $3\text{Na}_2\text{S}_4\text{O}_6 + 12\text{NaOH} = 3\text{Na}_2\text{S}_2\text{O}_3 + 5\text{Na}_2\text{SO}_3 + \text{Na}_2\text{S} + 6\text{H}_2\text{O}$. If cuprous chloride be dissolved in dilute hydrochloric acid containing some ammonium chloride, and a little thiosulphuric acid be added, a brown precipitate is formed which is a mixture of free sulphur, copper, and copper oxide. As indicated above, the reaction between iodine and sodium thiosulphate is much utilized in volumetric analysis for the quantitative determination of iodine.

Pentathionic acid, $\text{H}_2\text{S}_5\text{O}_6$.—When hydrogen sulphide is passed into a solution of sodium tetrathionate, the salt decomposes and nascent sulphur is probably formed: $\text{Na}_2\text{S}_4\text{O}_6 + 5\text{H}_2\text{S} = 2\text{NaOH} + 4\text{H}_2\text{O} + 9\text{S}$. The nascent sulphur then unites with the undecomposed tetrathionate, forming sodium pentathionate: $\text{Na}_2\text{S}_4\text{O}_6 + \text{S} = \text{Na}_2\text{S}_5\text{O}_6$. The same salt is formed, more or less contaminated with trithionate, when a solution of the tetrathionate decomposes: $2\text{Na}_2\text{S}_4\text{O}_6 = \text{Na}_2\text{S}_3\text{O}_6 + \text{Na}_2\text{S}_5\text{O}_6$; and it is also formed by the action of sulphur monochloride, S_2Cl_2 , on barium thiosulphate:



When hydrogen sulphide is passed into a concentrated solution of sulphurous acid at 0° , a solution containing a number of the polythionic acids is formed. It is called Wackenroder's solution. It is probable that the first action of the hydrogen sulphide results in the formation of tetrathionic acid: $\text{H}_2\text{S} + 3\text{SO}_2 = \text{H}_2\text{S}_4\text{O}_6$; and that this decomposes into the tri- and pentathionic acids as indicated above.¹ The

¹ According to H. Debus, part of the sulphur in Wackenroder's solution is a soluble variety which can be separated as an amorphous gelatinous mass which he called δ -sulphur. W. Spring considers this to be a hydrated form of sulphur, $\text{S}_8 \cdot n\text{H}_2\text{O}$, which can be separated from the accompanying polythionic acids by a few months' dialysis.

passage of the hydrogen sulphide can be continued until the solution contains little more than sulphur and pentathionic acid. The solution containing a mixture of sulphur and pentathionic acid can be concentrated by evaporation until its specific gravity is about 1.46, and then saturated with potassium hydroxide; the solution is filtered to remove the sulphur and allowed to crystallize spontaneously. The crystals which separate have the empirical composition: $K_2S_5O_6 \cdot 3H_2O$. The final products of the action are sulphur and water: $H_2S_5O_6 + 5H_2S = 6H_2O + 10S$; or $SO_2 + 2H_2S = 3S + 2H_2O$. The pentathionic acid, alone of the polythionic acids, almost immediately gives a brown precipitate with ammoniacal silver nitrate solutions. The precipitate soon becomes black. With potassium hydroxide, the pentathionates give an immediate precipitate of sulphur.

Potassium hexathionate, $K_2S_6O_6$, corresponding with the unknown hexathionic acid, $H_2S_6O_6$, is said to have been prepared in an impure condition from the mother liquid remaining after the separation of potassium pentathionate. The aqueous solution is very unstable. It reacts like the pentathionates with potassium hydroxide and ammoniacal silver nitrate, but it gives an *immediate* precipitate of sulphur with ammonia, the pentathionates give a precipitate with ammonia on standing a short time.

General properties.—Aqueous solutions of all the acids decompose when the attempt is made to concentrate them: $H_2S_nO_6 = H_2SO_4 + SO_2 + (n-2)S$; the solid salts decompose in a similar manner when heated. In dithionic acid, $n = 2$, and this corresponds with the fact that aqueous solutions of this acid yield no sulphur on decomposition, the other polythionic acids decompose with the separation of sulphur. Tetrathionic acid seems to be the most stable acid of the series. Unlike barium sulphate the barium salts of the polythionic acids are all soluble in water. The aqueous solutions of the alkaline polythionates decompose slowly. Mercurous nitrate gives no precipitate with dithionic acid, a black precipitate with trithionic acid and yellow precipitates with tetra- and pentathionic acids. Potassium hydroxide gives a precipitate of sulphur with pentathionic acid, and no precipitate with the other acids; dilute hydrochloric acid gives a precipitate of sulphur and evolution of sulphur dioxide with trithionic acid, and with the other acids no change is observed.

History.—Dithionic acid was discovered by J. J. Welter and J. L. Gay-Lussac in 1819; trithionic acid, by C. Langlois in 1842; tetrathionic acid, by M. J. Fordos and A. Gélis, 1843; and pentathionic acid by H. W. F. Wackenroder in 1845. The latter was specially studied by H. Debus in 1888. Thiosulphates were made by F. Chaussier in 1799, and afterwards carefully studied by L. N. Vauquelin in 1800.

§ 20. Review of the Oxides and Oxyacids of Sulphur.

We may now tabulate the list of the oxides and oxyacids of sulphur discussed in what precedes.

OXIDES.	ACIDS.
[Sulphur monoxide . . . SO]	[Sulphoxylic acid . . . H_2SO_2]
Sulphur sesquioxide . . . S_2O_3	Hyposulphurous acid . . . $H_2S_2O_4$
Sulphur dioxide . . . SO_2	Sulphurous acid . . . H_2SO_3
Sulphur trioxide . . . SO_3	Sulphuric acid . . . H_2SO_4
.	Pyrosulphuric acid . . . $H_2S_2O_7$

OXIDES.		ACIDS.	
Sulphur heptoxide . . .	S_2O_7	Persulphuric acid . . .	$H_2S_2O_8$
...	...	Caro's acid . . .	H_2SO_5
—	—	Thiosulphuric acid . . .	$H_2S_2O_3$
—	—	Dithionic acid . . .	$H_2S_2O_6$
—	—	Trithionic acid . . .	$H_2S_3O_6$
—	—	Tetrathionic acid . . .	$H_2S_4O_6$
—	—	Pentathionic acid . . .	$H_2S_5O_6$
—	—	[Hexathionic acid] . . .	$H_2S_6O_6$

Sulphoxylic acid, H_2SO_3 , is only known in the form of an organic derivative discussed in text-books of organic chemistry. Sulphur monoxide, SO , is unknown. The sulphones, R_2SO_2 , and the sulphinic acids, $R'SO.OH$, of organic chemistry, are related to the unknown sulphonylic acid, H_2SO_2 .

§ 21. Nascent Action.

At the moment of chemical change, the same chemical relation exists between the particles of which certain elements consists as between the particles of a compound substance under similar circumstances, on which relation the phenomena of combination depend; that in short (to use the common language) the particles of the elements have a chemical affinity for one another.—B. C. BRODIE (1850).

We have just alluded to the action of "nascent" sulphur; and some other examples of nascent action have been previously encountered, pp. 126, 282. Ordinary free hydrogen, oxygen, chlorine, etc., are unable to affect many substances which are readily attacked by mixtures known to yield hydrogen, oxygen, chlorine, etc. The term nascent action refers to the fact that an element at the moment of its separation appears to be more chemically active than after it has been made a few moments. Amongst the various hypotheses which have been suggested, three are plausible explanations of the phenomenon.

1. **Atomic hypothesis.**—Here it is assumed that nascent hydrogen is in the atomic condition and does its work before the atoms have had time to form ordinary molecules. There is, however, no direct evidence that, say, atomic hydrogen ever has a separate existence during a particular reaction (p. 344); and the facts can usually be explained on the energy hypothesis on the assumption that there is a double decomposition. Hence, E. Divers and T. Shimidzu (1885) said that "nascent hydrogen is an imaginary reagent."

2. **Ionic hypothesis.**—It will be remembered that the action of hydrochloric acid upon zinc, according to the ionic hypothesis, involves little more than the transfer of two positive charges from the hydrogen ion to the zinc atom: $Zn + 2H^+ + 2Cl^- = Zn^{++} + 2Cl^- + H_2$, and nascent hydrogen thus represents the condition of the element at the instant when its ions give up their electric charges.

3. **Energy hypothesis.**—During the reaction between, say, zinc and hydrochloric acid, energy is running down in the form of heat: $2Zn + 2HCl + Aq \rightarrow 2ZnCl_2 + H_2 + Aq + 112.8 \text{ Cals.}$ The greater activity of nascent hydrogen is ascribed to the energy of the reaction being available for inaugurating another reaction rather than being frittered away as heat (p. 344).

§ 22. Selenium and Tellurium.

The elements selenium and tellurium cannot be classed among the common elements. They are not very abundantly distributed in the "half-mile crust" of the earth. Small quantities of selenium are often found associated with sulphur and the sulphides, *e.g.* pyrites. Hence selenium is found in the "flue dust" of the "pyrites burners" in the manufacture of sulphuric acid. Small quantities of tellurium are found associated with gold, silver, and bismuth ores. Tellurium was recognized by the early mineralogists, and Muller von Reichenstein, in 1782, considered it to be a new element which M. H. Klaproth, in 1798, named tellurium—from the Latin *tellus*, the earth. Selenium was discovered by J. J. Berzelius, in 1817, and named selenium, from the Greek *σεληνή* (*selēnē*), the moon, owing to its resemblance to tellurium discovered a few years before.

Tellurium is a silver grey solid with a metallic lustre; selenium is a reddish-brown powder. Like sulphur, both elements exist in several allotropic forms, but the allotropism of tellurium is less marked than with selenium. Both elements conduct electricity, although one variety of selenium is almost a non-conductor. Both elements form hydrogen compounds— H_2Se and H_2Te —corresponding with hydrogen sulphide. These compounds are all prepared in a similar manner, and they respectively precipitate selenides and tellurides from solutions of salts of the metals. Hydrogen telluride is unstable even below 0° ; hydrogen selenide is rather more stable, but it decomposes in the light.

The two elements are monatomic at 2000° . Both elements burn in air with a blue flame, forming dioxides— SeO_2 and TeO_2 —analogous with sulphur dioxide. The dioxides dissolve in water, forming solutions analogous with sulphurous acid; and these, on oxidation, give selenic and telluric acids; which, in turn, form selenates and tellurates analogous with the sulphates. Selenic acid is weaker than sulphuric acid, and telluric acid is weaker than selenic acid. The two elements also form chlorides and bromides analogous with the corresponding sulphur chlorides, and bromides.

Selenium is used to a small extent in the production of certain violet and red colours for glass and enamels; and also for "bleaching," that

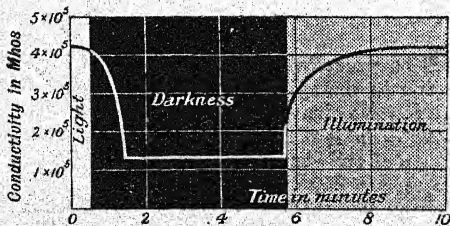
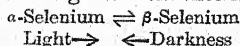


FIG. 186.—The Effect of Light and Darkness on the Electrical Conductivity of Selenium.

is, neutralizing, the green tinge of glass. Perhaps the most interesting property of selenium is the increase which occurs in its electrical conductivity or the decrease in its electrical resistance on exposure to light—discovered by W. Smith (1873). The change is proportional to the intensity of the light; red and orange light are

the most active. The phenomenon is illustrated by the curve, Fig. 186, where the electrical conductivity of selenium is plotted on passing from

light to darkness and back to light again. The change after exposure is not instantaneous, since in darkness the selenium takes a little time to recover its normal resistance. Siemens (1875) supposed that the phenomenon depended upon the change of selenium from a feeble conducting variety α -selenium, to a good conducting variety β -selenium; both forms assume a condition of equilibrium in light and in darkness, but in light the equilibrium is displaced in favour of the more conducting variety, and this the more the greater the intensity of the light:



The rate of change is accelerated by the presence of silver, platinum, etc., which act as catalytic agents. Many interesting applications have been made of this curious phenomenon—*e.g.* telephotography, wireless telephony, photometry, automatic control of the supply of gas in illuminated buoys, etc.

§ 23. The Oxygen Family of Elements.

Oxygen, sulphur, selenium, and tellurium form an interesting group of elements. The relationship is not so clearly defined as with the halogens. If we compare oxygen with tellurium, it would require some imagination to make the relationship significant, but on comparing any one element with its neighbour in the series: O, S, Se, Te, the relationship becomes more emphatic. They are all bivalent. The physical properties are best compared in the form of a table which brings out the gradation in properties very clearly. For instance, the specific gravities, melting points, boiling points, etc., increase with increasing atomic weight. Oxygen is at one end of the series, tellurium at the other:

TABLE XXX.—PROPERTIES OF THE OXYGEN-SULPHUR FAMILY.

	Oxygen.	Sulphur.	Selenium.	Tellurium.
Atomic weight	16	32.07	79.2	127.5
Melting point	-227°	114° – 115°	170° – 217°	452° – 454°
Boiling point	-183°	444.5°	688°	1390°
Specific gravity (solid) .	1.43	1.96–2.06	4.28–4.80	5.93–6.4
Atomic volume (approx.)	11	16	18	21
Colour of solid	Pale blue	Yellow	Reddish brown	Black
Heat of union with hydrogen (cals.)	69.0	4.8	-25.1	-34.9
State of aggregation of hydride	Liquid	Gas	Gas	Gas

Selenium, and tellurium particularly, lie very close to that ill-defined border line between the metals and non-metals. All four elements exhibit allotropism. Oxygen occurs most abundantly, sulphur next, and tellurium least. They all produce hydrogen compounds of the same type, but while the hydrogen compounds of sulphur, selenium, and tellurium are fetid-smelling gases at ordinary temperatures, hydrogen oxide, H_2O , is a colourless, odourless liquid. The ionization constants of the acids are:

Water.	Hydrosulphuric acid.	Hydroselenic acid.	Hydrotelluric acid.
10^{-14}	7.5×10^{-2}	1.7×10^{-4}	1.7×10^{-2}

The acidic character of the hydrides increases with increasing atomic weight. Tabulating the properties of the hydrides analogous with water we get:

TABLE XXXI.—PROPERTIES OF THE HYDRIDES OF THE SULPHUR-OXYGEN FAMILY.

Symbol.	Molecular weight.	Boiling point.	Melting point.	Specific gravity.	Dissociation temperature.	Reaction (litmus).
H ₂ O	18.02	100°	0°		1800°	Neutral
H ₂ S	34.09	-61.8°	-85.5°	1.17	400°	Acid
H ₂ Se	81.22	-42°	-64°	2.81	150°	Acid
H ₂ Te	129.52	0°	-48°	2.57	0°	Acid

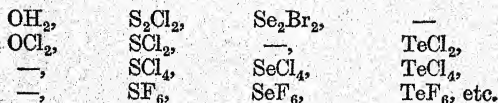
The first two also form H₂O₂ and H₂S₂ respectively. Sulphur, selenium, and tellurium unite with oxygen to form trioxides, but they do not form similar compounds with one another. The dioxides, however, form an interesting set if we regard ozone as an oxygen dioxide, thus:



Oxygen dioxide. Sulphur dioxide. Selenium dioxide. Tellurium dioxide.

Of course, the graphic formulæ could here have been given with oxygen, sulphur, etc., quadrivalent. The analogy between ozonic acid (HO)₂O : O and sulphurous acid (HO)₂S : O might also be cited.

Tellurium dioxide is an "intermediate oxide," because it acts both as an acid and as a base (p. 173)—thus, tellurium sulphate, Te(SO₄)₂, is known, and yet tellurous acid, H₂TeO₃, like sulphurous acid, forms salts called tellurites, *e.g.* K₂TeO₃, KHTeO₃, etc. The remarkable number of compounds of the metals with oxygen and sulphur, and the similarity in the constitution of the compounds of sulphur and of oxygen, has already been emphasized. Tellurium and selenium have not been so closely studied as sulphur and oxygen, but they undoubtedly show a similar behaviour in a less marked degree. All three form trioxides: SO₃, SeO₃, and TeO₃. All four elements form halides, although there are some gaps:



All four elements form compounds of the type of sulphides:



Questions.

1. Give a brief account of the manufacture of strong sulphuric acid. What are the three most common impurities in the concentrated acid? How may they be detected and removed?—*London Univ.*

2. Sulphuric acid is said to be dibasic, and to contain in its molecule two hydroxyl groups; on what evidence are these statements made? What is the origin of the name "oil of vitriol" by which this acid is known in commerce?—*London Univ.*

3. Explain exactly the meaning of the formula SO_2 . Describe how the composition of this gas may be determined.—*Aberdeen Univ.*

4. Explain the meaning of the italicized words: "The *anhydrous, neutral salt, insoluble in aqueous solvents*, was decomposed by an equivalent quantity of *dibasic acid*, producing some gaseous anhydride and a saturated solution of an inorganic salt without water of crystallization but very deliquescent."—*Dartmouth Coll., U.S.A.*

5. Give some examples of chemical changes which take place slowly. Have any observations been made between the rate and the condition of such change?—*New Zealand Univ.*

6. Describe the means you would adopt in order to prepare from sodium sulphite (a) sodium hyposulphite and (b) sodium thiosulphate. Discuss the existing evidence concerning the constitution of each of these three compounds.—*Science and Art Dept.*

7. Calculate the volume of a solution of sulphuric acid, density 1.8 and containing 89 per cent. of pure acid, that would be required to make 200 grams of hydrochloric acid by acting on sodium chloride. Atomic weights:— $\text{H}=1$, $\text{S}=32$, $\text{O}=16$, $\text{Na}=23$, $\text{Cl}=35.5$.—*Sheffield Scientific School, U.S.A.*

8. Why does sulphuric acid become hot when a limited amount of cold water is added to it? In what proportion is water added to produce the greatest heat?—*Amherst Coll., U.S.A.*

9. With sulphuric acid as the example, explain the meaning you attach to the terms "acid" and "dibasic acid." What is the behaviour towards litmus of NaHCO_3 , Na_2CO_3 , NaHSO_4 , Na_2SO_4 , and how do you account for it?—*Sheffield Univ.*

10.—(a) Why is sodium sulphate not utilized as a source of sulphuric acid, as sodium chloride is for hydrochloric acid?

(b) Describe briefly the manufacture of sulphuric acid by the "Contact Process."

(c) What is the specific gravity of concentrated commercial sulphuric acid? How much sulphuric acid does it contain? What is the present price per pound in carboy lots?—*Worcester Polytechnic Inst., U.S.A.*

11. Explain the reaction involved in the production of sulphuric acid in the "Chamber Process" and the means adopted to prevent the escape of nitrous gases into the atmosphere. Give two illustrations of the use of sulphuric acid as (a) a dehydrating agent, (b) an oxidizing agent.—*Sheffield Univ.*

12. In chemical literature at the present day reference is frequently made to "condensing agents," and substances are spoken of as "undergoing condensation." Explain these phrases, giving examples in illustration of your answer.—*London Univ.*

13. Chlorine is said to be an oxidizing agent, and sulphur dioxide a reducing agent; explain these statements and give examples illustrating these actions, stating any conditions which may be necessary. Give also examples of reactions in which sulphur acts as a reducing agent.—*London Univ.*

14. What happens on heating with sulphuric acid (a) manganese dioxide, (b) cuprous oxide; with nitric acid (a) black oxide of iron, (b) red lead; with hydrochloric acid (a) red oxide of iron, (b) black oxide of manganese?—*London Univ.*

15. The composition of a compound is: barium 46.12 per cent., sulphur 21.54 per cent., and oxygen 32.32 per cent. What is the formula and name of the compound?—*Glasgow Univ.*

16. Explain the meaning of the term "nascent" in chemistry, illustrating your answer by reference to reactions in which (a) "nascent" hydrogen, (b) "nascent" oxygen are supposed to take part.—*Board of Educ.*

17. Discuss the acids formed by the solution of SO_2 and SO_3 in water. What is a sulphonic acid?—*Owens Coll.*

18. Give graphic formulæ for sulphurous and sulphuric acids, with a clear account of the facts which have led to their adoption. What happens when a solution of sulphur dioxide is treated with zinc?—*Board of Educ.*

19. How is sodium thiosulphate made? What is the meaning of the name "thiosulphate"? Why is the name "hyposulphite" often applied to this salt inappropriately? Describe the true hyposulphites.—*Board of Educ.*

20. A metal forms only one series of salts and the anhydrous sulphate contains 58.8 per cent. of the metal. Calculate the equivalent of the element; and the percentage composition of the anhydrous chloride.—*Customs and Excise.*

21. Certain substances take up moisture when exposed to the air; certain other substances tend to lose water of crystallization under similar conditions. Give a full statement of the principle underlying these facts, and define the two terms ordinarily used to describe them.—*Cornell Univ., U.S.A.*

22. What is the experimental basis of structural formulæ? Explain the proper significance of such formulæ. Should they be understood to imply any actual knowledge of the structure of molecules, or are they merely "condensed equations"?—*New Zealand Univ.*

23. Describe briefly the qualitative and quantitative experiments you would consider necessary to show that the formulæ usually adopted for potassium persulphate is correct.—*Calcutta Univ.*

24. What processes are involved in the setting of ordinary plaster, and of plaster of Paris? How would you convert plaster of Paris into bleaching powder?—*Sydney Univ.*

25. What is gypsum? Describe the preparation of plaster of Paris. Explain the change which causes the setting of the plaster.—*Science and Art Dept.*

26. Write the formulæ of the chief alums known. Point out their character as a class. State and explain the principle they illustrate.—*London Univ.*

27. What is meant by an alum? How can the fact that chromium forms an alum be utilized in fixing its atomic weight?—*London Univ.*

28. Describe the preparation, properties, and uses of plaster of Paris. Compare the setting of plaster of Paris, air mortar, and cement.—*Cape Univ.*

29. Give an account of the phenomenon of catalysis. Describe any one manufacturing process which depends for its operation on the action of a catalyser.—*Board of Educ.*

30. Describe fully an experiment by means of which it may be shown that sulphur dioxide gas contains its own volume of oxygen. Explain clearly the reactions, if any, which take place when this gas is passed into each of the following liquids:—(a) lime water, (b) caustic soda solution, (c) an acidified solution of potassium permanganate, (d) potassium iodate solution, (e) fuming nitric acid.—*Board of Educ.*

31. Describe the means you would adopt in order to prepare from sodium sulphite (a) sodium hyposulphite and (b) sodium thiosulphate. Discuss the existing evidence concerning the constitution of each of these three compounds.—*Science and Art Dept.*

32. Describe the changes which occur when oil of vitriol is heated with copper. State any reasons that may be given for or against the theory that the so-called "nascent hydrogen" is concerned on this reaction.—*London Univ.*

33. Plot the following results for the specific gravities of solutions of sulphuric acid of different concentrations:

Grams H_2SO_4 per litre . . .	16	257	510	1015	1404	1621	1825
Specific gravity	1.010	1.180	1.300	1.560	1.740	1.815	1.840

(a) Indicate the weight H_2SO_4 in a litre of acid of specific gravity 1.25, and (b) state the percentage amount of H_2SO_4 in an acid of specific gravity 1.3. (Answers: (a) 418; (b) 39.19.)

CHAPTER XXIV

CHROMIUM, MOLYBDENUM, TUNGSTEN, AND URANIUM

§ 1. Potassium and Sodium Dichromates.

Chromite.—This mineral is also called chrome ironstone, or chrome iron ore. It is a native ferrous chromite, FeCr_2O_4 , or $\text{Fe}(\text{CrO}_2)_2$, contaminated with silica, alumina, magnesia, etc. It resembles magnetite in general appearance, for it has an iron-black colour, with a brownish tinge. Good marketable chromite contains the equivalent of at least 50 per cent. Cr_2O_3 , and not more than about 10 per cent. of silica. Most of the chromium compounds of commerce are derived from this ore.

Manufacture of sodium chromate.—The finely ground chromite is intimately mixed with lime and sodium carbonate, and roasted in an oxidizing atmosphere. The reaction which occurs is probably: $4\text{FeCr}_2\text{O}_4 + 8\text{Na}_2\text{CO}_3 + 7\text{O}_2 = 8\text{Na}_2\text{CrO}_4 + 8\text{CO}_2 + 2\text{Fe}_2\text{O}_3$. Possibly the ferrous chromite produces sodium chromite, NaCrO_2 , thus: $\text{Fe}(\text{CrO}_2)_2 + \text{Na}_2\text{CO}_3 = 2\text{NaCrO}_2 + \text{FeO} + \text{CO}_2$. The ferrous oxide and the sodium chromite are then oxidized by the air. The object of the lime is to prevent fusion, and keep the mass porous to facilitate oxidation. The roasted mass is then mixed with twice its weight of water, and an excess of soda ash is added to convert the calcium chromate formed during the reaction into sodium chromate. The mixture is agitated for a couple of hours, and the solution of sodium chromate is separated from the insoluble matters by filter presses.

Transformation of sodium chromate to the dichromate.—The clear solution is then treated with sulphuric acid to neutralize the excess of alkali, and to convert the sodium chromate, Na_2CrO_4 , into sodium dichromate, $\text{Na}_2\text{Cr}_2\text{O}_7$, thus: $\text{H}_2\text{SO}_4 + 2\text{Na}_2\text{CrO}_4 = \text{H}_2\text{O} + \text{Na}_2\text{SO}_4 + \text{Na}_2\text{Cr}_2\text{O}_7$. The clear solution is decanted from the precipitated sodium sulphate, and the solution is separated from the crystals by centrifugal separators. The solution is concentrated in iron pans, and when it has attained the specific gravity 1.7, it is filtered from the sodium sulphate which has separated during the evaporation. Crystals of sodium dichromate separate on standing. The yield is about 90 per cent. of the theoretical.

Conversion of sodium dichromate into the potassium salt.—The sodium dichromate is converted into the potassium salt by mixing concentrated solutions of sodium dichromate with potassium chloride: $\text{Na}_2\text{Cr}_2\text{O}_7 + 2\text{KCl} \rightleftharpoons \text{K}_2\text{Cr}_2\text{O}_7 + 2\text{NaCl}$. The potassium dichromate is at once precipitated. The chromium is recovered from the mother liquid. The chromate can be converted into the dichromate electrolytically by using it as catholyte in a cell with a porous partition using sodium hydroxide as anolyte. This apparently roundabout method of making sodium

dichromate in order to get the potassium salt is really cheaper than if potassium carbonate had been used in place of sodium carbonate, mainly because of the difference in cost between potassium carbonate, and sodium carbonate and potassium chloride.

Properties of sodium and potassium dichromates.—Potassium dichromate forms bright red triclinic crystals of the anhydrous salt, $K_2Cr_2O_7$, which melt at 400° . The salt is easily purified by recrystallization because it is very much more soluble in hot than in cold water: 100 grams of water dissolve 5 grams of the salt at 0° , 55.5 grams at 50° , and 129 grams at 100° . The aqueous solution has an acid reaction. The fact that potassium dichromate is so easily purified makes it a good starting-point for the manufacture of chromium compounds generally. Sodium dichromate crystals have the composition $Na_2Cr_2O_7 \cdot 2H_2O$. This salt is cheaper and more soluble in water than the potassium salt at ordinary temperatures. At 0° , 100 grams of the solution contain 62 grams of $Na_2Cr_2O_7$; and at 80° , 80 grams of the salt. A monoclinic variety of potassium dichromate also crystallizes from a hot concentrated solution of the salt containing a little potassium thiocyanate. The monoclinic variety is unstable and soon passes into the triclinic form. The specific gravity of the triclinic form is 2.67, and of the monoclinic form, 2.10.

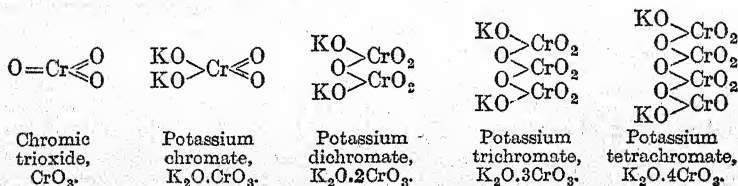
§ 2. Chromic Acid and the Polychromates.

Chromium trioxide, CrO_3 .—When a concentrated aqueous solution of potassium dichromate is treated with concentrated sulphuric acid, long scarlet needle-like crystals separate when the solution is cooled. The crystals can be filtered through glass wool and drained on a porous earthenware tile; then washed with concentrated nitric acid to remove the sulphuric acid and sulphates, and dried in a current of dry warm air. The resulting crystals of chromium trioxide, CrO_3 , are very deliquescent, and dissolve in water, forming a red solution which is probably dichromic acid, $H_2Cr_2O_7$, formed by the reaction: $2CrO_3 + H_2O = H_2Cr_2O_7$. The acid H_2CrO_4 has not been isolated. Chromic trioxide, CrO_3 , is thus called chromic anhydride and the aqueous solution chromic acid, although the trioxide crystallizes out again when the aqueous solution is evaporated. Chromium trioxide can be volatilized just over 110° , and it melts at about 196° , with some decomposition, and it soon decomposes when heated to about 250° into chromic oxide Cr_2O_3 , and oxygen: $4CrO_3 = 2Cr_2O_3 + 3O_2$. Chromium trioxide is a vigorous oxidizing agent owing to the readiness with which it seems to part with oxygen to form chromic oxide, Cr_2O_3 . Thus, alcohol dropped on to the oxide takes fire; when ammonia gas impinges on the crystals, the reduction takes place with incandescence, paper is charred at once; carbonaceous matter is oxidized to carbon dioxide; etc.

Potassium chromate, K_2CrO_4 .—If potassium hydroxide be added to an aqueous solution of potassium dichromate in the proportions indicated by the equation: $K_2Cr_2O_7 + 2KOH = H_2O + 2K_2CrO_4$, a yellow solution of potassium chromate is obtained which furnishes bright yellow rhombic crystals when concentrated. The yellow crystals of potassium chromate, K_2CrO_4 , are isomorphous with potassium sulphate. Hence potassium chromate may be regarded as a salt formed by the union of one molecule

of potash, K_2O , with one molecule of chromic anhydride, CrO_3 , thus, $K_2O \cdot CrO_3$ or K_2CrO_4 . The isomorphism with potassium sulphate suggests a similar structural formula with an atom of sexivalent chromium in place of sexivalent sulphur. If normal chromates be treated with acids dichromates are formed: $2K_2CrO_4 + 2HCl = 2KCl + K_2Cr_2O_7 + H_2O$. Potassium dichromate is a salt formed by the union of two molecules of the anhydride, CrO_3 , with one molecule of the base: $K_2O \cdot 2CrO_3$.

Polychromates.—The radicle CrO_4^{--} is bivalent in the chromates, and the radicle $Cr_2O_7^{--}$ is bivalent in the dichromates. By treating potassium dichromate with chromic oxide, or with boiling moderately concentrated nitric acid, potassium trichromate is formed: $K_2Cr_3O_{10}$, or $K_2O \cdot 3CrO_3$; and by treating the trichromate with concentrated nitric acid, potassium tetrachromate, $K_2O \cdot 4CrO_3$, or $K_2Cr_4O_{13}$, is formed. Just as in the formation of disulphuric or pyrosulphuric acid (*q.v.*) and its salts by the condensation of two molecules of H_2SO_4 , so here, dichromic acid is considered to be a condensation product of two molecules of chromic acid, H_2CrO_4 ; and the dichromates, *e.g.* $K_2Cr_2O_7$, are analogous with the pyrosulphates, *e.g.* $K_2S_2O_7$. The constitutional formulæ of chromic trioxide is supposed to be analogous with that of sulphur trioxide, sulphur and chromium being both sexivalent; the chromates and dichromates are also supposed to be analogous with the constitutional formulæ of the sulphates and pyro- or disulphates respectively:



The addition of an excess of alkali to the polychromates reconverts them into normal chromates; and water converts them into dichromates. The tetrachromate is fairly stable in the presence of an excess of chromic trioxide. As the relative concentrations of chromic acid and potassium hydroxide in aqueous solution is increased from zero chromic acid upwards, the compounds which separate successively in the solid state are the hydroxide, chromate, dichromate, trichromate, tetrachromate, and finally chromic acid. The phenomenon of condensation also occurs with many oxyacids—boric, iodic, and phosphoric acids—and particularly with molybdic and tungstic acids; it is slight with uranic acid, and unknown with nitric acid.

Chromates and dichromates.—Ammonium dichromate, $(NH_4)_2Cr_2O_7$, is formed when aqueous solutions of ammonia and of chromic acid are mixed. When the solid is heated, free nitrogen, water, and a voluminous mass of chromic oxide (whose appearance has been likened to green tea) are obtained: $(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + 4H_2O + N_2$. This is an interesting but costly way of making nitrogen. The percentage solubility of the alkali chromates and dichromates at 30° are:

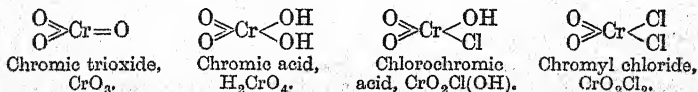
	Li	Na	K	Rb	NH_4
Chromate . . .	50.0	46.6	39.4	44.1	28.8
Dichromate . .	56.6	66.4	15.3	9.5	32.1

All the chromates of lithium and sodium crystallize with water, but the chromates of potassium, sodium, rubidium, and ammonium are anhydrous. Lead chromate, PbCrO_4 , is a bright sulphur-yellow salt precipitated by adding potassium chromate to the solution of a lead salt. It is used as a pigment under various trade-names—chrome yellow, Cologne yellow, lemon chrome, etc. By boiling lead chromate with aqueous ammonia or potassium hydroxide, a basic lead chromate, $\text{Pb}(\text{OH})_2\text{PbCrO}_4$, called chrome red, Austrian cinnabar, American vermilion, Parisian red, etc., is formed. This also is used as a pigment. In the formation of this basic salt, the stronger base sodium displaces part of the lead from lead chromate: $2\text{PbCrO}_4 + 2\text{NaOH} \rightleftharpoons \text{PbO} \cdot \text{PbCrO}_4 + \text{Na}_2\text{CrO}_4 + \text{H}_2\text{O}$; with sodium carbonate and lead chromate: $2\text{PbCrO}_4 + 2\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons 2\text{NaHCO}_3 + \text{Na}_2\text{CrO}_4 + \text{PbO} \cdot \text{PbCrO}_4$. The same equilibrium state is not obtained when the start is made from the lead chromate side as from the basic chromate side. Barium chromate, BaCrO_4 , is formed in a similar manner, and is used as a pigment. The precipitation of barium chromate from a solution of a barium salt by adding a soluble chromate; or conversely, the precipitation of chromium as barium chromate by adding a soluble barium salt, enables the amount of barium or of chromium in a solution to be determined. At 18° a litre of water only dissolves 0.0038 gram of barium chromate; 1.2 gram of strontium chromate; and 23.2 grams of calcium chromate. Calcium chromate, $\text{CaCrO}_4 \cdot 2\text{H}_2\text{O}$, is isomorphous with calcium sulphate, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$. Silver chromate and mercurous chromate are both red salts. The chromates are often made by the addition of soluble dichromates to a solution of the salt in question. The ionic hypothesis describes the reaction thus: The dichromate ions, $\text{Cr}_2\text{O}_7^{2-}$, in the solution are partly broken down into CrO_4^{2-} ions. For equilibrium: $\text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O} \rightleftharpoons 2\text{CrO}_4^{2-} + 2\text{H}^+$. If any Pb^{2+} ions are present, the CrO_4^{2-} ions are removed because PbCrO_4 separates from the solution. The supply of CrO_4^{2-} ions is kept up by the continued dissociation of the $\text{Cr}_2\text{O}_7^{2-}$ ions until all the dichromate has been converted into chromate.

Potassium dichromate as an oxidizing agent.—Potassium dichromate is used in volumetric analysis in virtue of its oxidizing properties when in contact with a reducing agent, e.g. ferrous sulphate, FeSO_4 . Since potassium dichromate has a formula equivalent to $\text{K}_2\text{O} \cdot 2\text{CrO}_3$, and ferrous sulphate a formula equivalent to $\text{FeO} \cdot \text{SO}_3$, the 2CrO_3 of the former on reduction furnish $2\text{CrO}_3 = \text{Cr}_2\text{O}_3 + 3\text{O}$; and the ferrous oxide, FeO , of the latter is converted into Fe_2O_3 , it follows that one molecule of potassium dichromate is equivalent in oxidizing properties to three atoms of oxygen, and it can therefore oxidize six molecules of ferrous sulphate. Hence the equation can be written: $6\text{FeSO}_4 + \text{K}_2\text{Cr}_2\text{O}_7 = 3\text{Fe}_2\text{O}_3 + \text{Cr}_2\text{O}_3 + 6\text{SO}_3 + \text{K}_2\text{O}$. The action takes place in an acidified solution so that the ferrous sulphate is oxidized to ferric sulphate, $\text{Fe}_2(\text{SO}_4)_3$, the potassium oxide forms potassium sulphate, K_2SO_4 , and the chromic oxide, chromic sulphate, $\text{Cr}_2(\text{SO}_4)_3$. In all, thirteen SO_4 radicles are needed, but six SO_4 radicles already come from the ferrous sulphate, hence seven molecules of sulphuric acid are needed. The full equation thus becomes: $6\text{FeSO}_4 + \text{K}_2\text{Cr}_2\text{O}_7 + 7\text{H}_2\text{SO}_4 = 3\text{Fe}_2(\text{SO}_4)_3 + \text{Cr}_2(\text{SO}_4)_3 + \text{K}_2\text{SO}_4 + 7\text{H}_2\text{O}$. Similarly, one molecule of potassium dichromate can oxidize three molecules of sulphurous acid, H_2SO_3 , to sulphuric acid, H_2SO_4 , and the equation is accordingly written: $\text{K}_2\text{Cr}_2\text{O}_7 + 3\text{H}_2\text{SO}_3$

+ $\text{H}_2\text{SO}_4 = \text{Cr}_2(\text{SO}_4)_3 + \text{K}_2\text{SO}_4 + 4\text{H}_2\text{O}$. When heated with hydrochloric acid, chlorine is produced as indicated in our study of the action of oxidizing agents on hydrochloric acid.

Chromyl chloride, CrO_2Cl_2 .—When potassium dichromate, a soluble chloride, and sulphuric acid are heated in a retort, a dark reddish-brown liquid, which boils at 116° , distills over. It is chromic oxychloride, or chromyl chloride, CrO_2Cl_2 . The distillation is stopped when the frothing ceases or the distillate will be contaminated with decomposition products. Chromyl chloride is decomposed into chromic acid and hydrochloric acid by contact with water: $\text{CrO}_2\text{Cl}_2 + 2\text{H}_2\text{O} = 2\text{HCl} + \text{H}_2\text{CrO}_4$. These reactions are used as a test for chlorides. Neither the bromide nor the iodide form corresponding compounds, so that if a mixture of these three halogens be distilled with sulphuric acid and potassium dichromate, the distillate, when treated with water, gives a solution which responds to the tests for chromic acid, the presence of chlorides may be inferred. Chromyl chloride is also formed by dissolving chromic trioxide in concentrated sulphuric acid, adding hydrochloric acid, drop by drop, and distilling the mixture as before: $\text{CrO}_3 + 2\text{HCl} = \text{CrO}_2\text{Cl}_2 + \text{H}_2\text{O}$. The sulphuric acid retains the water formed during the reaction. Yellowish-red crystals of potassium chlorochromate, $\text{CrO}_2\text{Cl}(\text{OK})$, corresponding with the unknown chlorochromic acid, $\text{CrO}_2\text{Cl}(\text{OH})$, are formed when a solution of potassium dichromate in hydrochloric acid is allowed to crystallize: $\text{K}_2\text{Cr}_2\text{O}_7 + 2\text{HCl} = \text{H}_2\text{O} + 2\text{CrO}_2(\text{OK})\text{Cl}$. The two chlorides are supposed to be related to one another as follows:

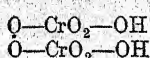


Chromyl chloride thus appears to be analogous with sulphuryl chloride, SO_2Cl_2 (*q.v.*), and chlorochromic acid with chlorosulphonic acid, $\text{SO}_2\text{Cl}(\text{OH})$. The molecular weight of chromyl chloride in freezing benzene is normal, but in boiling benzene it is more than doubled. The vapour density is 154. Chromyl fluoride, CrO_2F_2 , is known, but not the corresponding bromide and iodide. The chloride corresponding with chromium trioxide, namely, CrCl_6 , has not been prepared. The evidence for chromyl fluoride, CrO_2F_2 , is ambiguous, because the reddish gas formed when a mixture of calcium fluoride, sulphuric acid, and potassium dichromate is heated has not been proved to have this composition.

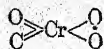
Perchromic acid.—In 1847 Barreswell discovered that a deep blue solution is obtained by treating dilute solutions of chromic acid, or acidified chromates, with hydrogen peroxide. This reaction is common use as a means of detecting both chromic acid and hydrogen peroxide (*q.v.*). By shaking the mixture up with ether, an ethereal solution of the blue compound formed during the reaction can be obtained. By treating the blue solution with potassium at a temperature below -20° , hydrogen is copiously evolved, and a dark purple precipitate with the empirical composition KCrO_4 separates. The molecular formula is supposed to be $\text{K}_2\text{Cr}_2\text{O}_8$ or potassium perdichromate analogous with the perdisulphates—sulphur and chromium sexivalent. This salt decomposes at ordinary temperatures into potassium dichromate and oxygen: $2\text{K}_2\text{Cr}_2\text{O}_8 = \text{O}_2 + 2\text{K}_2\text{Cr}_2\text{O}_7$. The blue solution must contain perchromic acid corresponding

to this salt. Lithium, sodium, magnesium, calcium, barium, and zinc perchromates have been prepared by the action of the acetates of these elements upon the blue solution. Ammonia gas, at -40° , yields the ammonium salt.

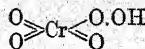
If an excess of hydrogen peroxide be employed in the preparation of the blue solution, stable higher chromates have been produced. Thus by adding a 30 per cent. solution of hydrogen peroxide to an alcoholic solution of potassium perchromate, a dark red precipitate of another potassium perchromate, K_3CrO_8 , is obtained. This appears to be stable below 70° , and is considered to be a salt of an unknown acid, H_3CrO_8 . Riesenfeld and Man (1914) obtained a blue crystalline precipitate by treating with hydrogen peroxide a solution of chromic anhydride in methyl ether at -30° . The composition agreed with the hypermonochromate, $H_3CrO_8 \cdot 2H_2O$, and it was suggested that the water is constitutional, not merely water of crystallization. The chromium is supposed to be septivalent, $(HO)_4Cr(O.OH)_3$; and the anhydride, $O_3Cr(O.OH)_3$. The composition of the acid did not vary with the amount of hydrogen peroxide used to produce it. At temperatures higher than -30° it decomposed, and this explains the failure of previous attempts to isolate it. Its strength as an acid was estimated to be about the same as that of acetic acid. The perchromates have been divided into two classes, (i) derivatives of chromium tetroxide, and (ii) the perchromates proper. In the former, chromium is sexivalent, in the latter septivalent.



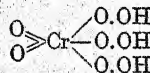
Perdichromic
acid, $H_2Cr_2O_8$.



Chromium
tetroxide, CrO_4 .



Permonochromic
acid, $HCrO_5$.



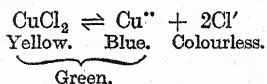
Hypermonoper-
chromic acid, H_3CrO_8 .

When a cold sulphuric acid solution of potassium dichromate is treated with an excess of hydrogen peroxide, the violet product corresponds with a permonochromate, $KCrO_5$, which decomposes in aqueous solution. If the mixing be in the reverse order, the hypermonochromate, K_3CrO_8 , is formed. When chromic acid is treated with hydrogen peroxide in the presence of ammonia or potassium cyanide, chromic triamminotetroxide, $Cr(NH_3)_3O_4$, is formed. It is supposed that the fundamental oxide is chromium tetroxide, CrO_4 , in which the chromium is sexivalent. We are, however, not very clear about even the empirical composition of these higher chromates; the methods of preparing pure salts are not satisfactory.

§ 3. The Colours of Salt Solutions.

According to the ionic hypothesis, the colour of a dilute aqueous solution of an electrolyte is an additive effect of the colours of the anions, the cations, and of the unionized molecule. The colour of the latter may be quite different from the colours of the two former so that the colour changes as the solution is more and more diluted, until ionization is complete. The action of water on cupric chloride or cupric bromide illustrates the idea very well. Solid cupric chloride, $CuCl_2$, is a dark brown powder which, when treated with a very small quantity of water, gives a yellow solution. This is supposed to represent the colour of the molecules

CuCl_2 . When the solution is still further diluted the colour becomes green, and finally blue. The blue colour is supposed to represent the colour of the Cu^{++} ions; the Cl^- ions are supposed to be colourless. The green colour is due to the partial ionization of the salt, and the sensation of green is due to the mixing of the yellow colour of the CuCl molecules with the blue colour of the Cu^{++} ions—



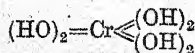
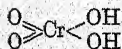
If a concentrated solution of ammonium chloride, hydrochloric acid, or a dehydrating agent—aluminium chloride, calcium chloride, etc.—be added to the blue solution, not too dilute, the ionization is supposed to be driven back, and blue cupric ions suppressed, as indicated on p. 379, because the solution becomes green. Feebly ionized chlorides, *e.g.* mercuric chloride, do not restore the green colour. Other copper salts give similar results. Dilute solutions of equivalent concentration give a similar colour in spite of the fact that the salts are different. Hence it is assumed that the Cu^{++} ions are coloured blue; and that the ions behave with respect to colour independently of one another.

Since cobaltous chloride gives a pink colour in dilute solutions, it is inferred that Co^{++} ions are this colour. If concentrated hydrochloric acid be added, the solution becomes blue owing to the formation of blue cobaltous chloride molecules. The same effect is produced by raising the temperature, the pink colour becomes blue because it follows that the degree of ionization is decreased by raising the temperature. The true explanation in the case of cobaltous chloride is probably not so simple as this, because while calcium chloride, CaCl_2 , turns the red solution blue, zinc chloride, ZnCl_2 , turns the blue solution red. This is supposed to be due to the formation of complex salts, CaCoCl_4 , which gives blue CoCl_4^{--} ions; and $\text{Co}(\text{ZnCl}_4)$, which yields colourless ZnCl_4^{--} ions and red Co^{++} ions. Many other hypotheses have been suggested—hydration and dehydration, isomerism, etc.

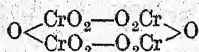
Among the coloured ions Cu^{++} is blue; Fe^{++} , Ni^{++} , Cr^{+++} , MnO_4^- are green; Co^{++} red; Mn^{++} pale pink; MnO_4^- purple; CrO_4^{--} yellow; $\text{Cr}_2\text{O}_7^{--}$ orange red; etc. The colourless ions include Cl^- , I^- , Br^- ; CyS^- , NO_3^- ; PO_4^{--} ; ClO_3^- ; K^+ ; Na^+ ; Ca^{++} ; Mg^{++} ; Fe^{+++} ; Pb^{++} ; SO_4^{--} ; etc. The colours of the ions are deduced, as indicated above, from the effects of dilution on the colour of the aqueous solutions.

The ionic hypothesis assumes that the difference in the colour of aqueous solutions of potassium chromate and potassium dichromate is due to the difference in the colours of the CrO_4^{--} ions of the chromates, and the $\text{Cr}_2\text{O}_7^{--}$ ions of the dichromates; the former are yellow, the latter orange or red. The CrO_4^{--} ions are supposed to be unstable in the presence of the H^+ ions of acids: $2\text{CrO}_4^{--} + 2\text{H}^+ = \text{Cr}_2\text{O}_7^{--} + \text{H}_2\text{O}$; and the dichromate ions unstable in the presence of OH^- ions of alkaline solutions: $\text{Cr}_2\text{O}_7^{--} + 2\text{OH}^- = 2\text{CrO}_4^{--} + \text{H}_2\text{O}$. Since chromium trioxide, CrO_3 , gives an orange coloured solution with water, it is inferred that dichromic, not chromic, acid is formed when CrO_3 is dissolved in water. This view is confirmed by observations on the depression of the freezing point of aqueous solutions of chromium trioxide; from measurements on the electrical conductivity

of the aqueous solutions; from a study of the colour changes on dilution; absorption spectra; etc. The general result is to show that very dilute solutions of the chromates contain only chromate ions; with more concentrated solutions both chromate and dichromate ions are present, so that chromate solutions contain dichromates and *vice versa*; and at still greater concentrations trichromates are formed, and tetrachromates at still greater concentrations. In general symbols: $\text{H}_2\text{CrO}_4 \rightleftharpoons \text{H}_2\text{Cr}_2\text{O}_7 \rightleftharpoons \text{H}_2\text{Cr}_3\text{O}_{10} \rightleftharpoons \text{H}_2\text{Cr}_4\text{O}_{13}$. The detailed equations would represent these changes as hydrolytic equilibria: $2\text{H}_2\text{CrO}_4 \rightleftharpoons \text{H}_2\text{O} + \text{H}_2\text{Cr}_2\text{O}_7$; $\text{H}_2\text{Cr}_3\text{O}_{10} + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{CrO}_4$; $\text{H}_2\text{Cr}_4\text{O}_{13} + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{Cr}_3\text{O}_{10} + \text{H}_2\text{CrO}_4$; etc. The passage from left to right is favoured by increasing concentration, heat, and acids; and conversely, the right to left transformations are favoured by dilution, cold, and the presence of alkalis. At extreme dilutions, when the dichromate is all hydrolyzed, it is probable that the monochromic acid H_2CrO_4 is hydrated first: $\text{H}_2\text{CrO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{H}_4\text{CrO}_5$; and then: $\text{H}_2\text{O} + \text{H}_4\text{CrO}_5 \rightleftharpoons \text{H}_6\text{CrO}_6$, because (1) the quality of the colour changes in solutions of less concentration than 0.01 per cent.; (2) the sensitiveness of the colour exhibits two maximal points; and (3) several of the chromates crystallize with combined water—e.g. $\text{Na}_2\text{CrO}_4 \cdot 10\text{H}_2\text{O}$; $\text{MgCrO}_4 \cdot 7\text{H}_2\text{O}$, etc. Graphically, these hydrated forms of monochromic acid are represented:



It is further supposed that since the tetrachromate is formed in *very* concentrated solutions, and that chromic trioxide is the anhydrous form, chromic trioxide is probably a polymer, $(\text{CrO}_3)_4$, a kind of ring compound,



When chromic anhydride is hydrolyzed to tetrachromic acid, the ring is broken to form the chain: $\text{HO}-\text{CrO}_2-\text{O}-\text{CrO}_2-\text{O}-\text{CrO}_2-\text{O}-\text{CrO}_2-\text{OH}$; and this in turn may be hydrolyzed to form one or more of the lower chromic acids when water enters one or more of the alternate oxygen linkages (dotted lines).

§ 4. Chromium Oxides and Hydroxides.

Chromium hydroxide, $\text{Cr}(\text{OH})_3$.—By adding potassium hydroxide to a solution of chromous chloride (*q.v.*) a yellowish brown precipitate of chromous hydroxide is obtained which rapidly oxidizes in air. The corresponding chromous oxide, CrO , has been obtained as a black powder by warming chromium amalgam with dilute nitric acid. The oxide is insoluble in this acid, and in dilute sulphuric acid, but is soluble in dilute hydrochloric acid with the evolution of hydrogen. It is reduced by hydrogen at 1000° , whereas chromic oxide is not affected by hydrogen at 1300° .

Chromium sesquioxide, Cr_2O_3 .—This oxide is prepared as a dark green powder when ammonium dichromate is heated, or when a mixture of potassium dichromate and ammonium chloride is heated. In the latter case, the potassium chloride is removed by washing the residue with water.

The oxide calcined at a high temperature is dissolved by acids very slowly. Chromic hydroxide, $\text{Cr}(\text{OH})_3$, separates as a bluish gelatinous (colloidal) precipitate when ammonia is added to a solution of a chromic salt. Chromic hydroxide dissolves in a solution of chromic chloride, and if the solution be dialyzed, as in the case of ferric chloride, a colloidal solution, hydrosol of chromic hydroxide, is obtained. The freshly prepared hydroxide can absorb carbon dioxide from the atmosphere. When freshly precipitated, chromic hydroxide readily dissolves in acids, but it is less rapidly dissolved if it has stood some time. When heated in air, it forms green chromic oxide, Cr_2O_3 . Several important green pigments are made by preparing the oxide under special conditions, *e.g.* calcining potassium dichromate intimately mixed with sulphur, ammonium chloride, starch, boric acid, etc., and extracting the matters soluble in water. Guignet's green, for instance, is made by calcining potassium dichromate with boric acid, etc. Chromic sesquioxide, in an extremely fine state of subdivision, appears to be crimson, for if an intimate mixture of stannic oxide, or zinc oxide, or alumina, with a very small proportion of chromic oxide, be heated to a high temperature in an oxidizing atmosphere a red powder is obtained. There is some evidence to show that the red colour is not due to the formation of a chemical compound, and that the "chrome-tin" colour is related to purple of Cassius (*q.v.*). The "chrome-tin" crimson is used for colouring pottery, glazes, etc. The chrome-alumina colour can be prepared to appear green in daylight or in reflected light, and crimson in transmitted or in artificial light, thus resembling the mineral *alexandrite*. Chromium is the colouring agent of artificial rubies, first synthesized in 1837 by A. Gaudin by melting together potassium dichromate and alumina in the oxyhydrogen blowpipe.

Chromic hydroxide is a base, and forms salts—chromic chloride, sulphate, etc.—when treated with the proper acid. It is also a feeble acid, for, when freshly precipitated, it dissolves in alkali hydroxides presumably owing to the formation of alkaline chromites—*e.g.* $\text{Cr}(\text{OH})_2\text{OK}$, or $\text{CrO}.\text{OK}$, that is, KCrO_2 . Native chromite is a ferrous chromite, $\text{Fe}(\text{CrO}_2)_2$, its constitution is probably analogous with the spinels (*q.v.*). The chromites are regarded as derivatives of an unknown chromous acid, HCrO_2 . The soluble chromites are hydrolyzed when their aqueous solutions are boiled, and greenish chromic hydroxide is precipitated.

§ 5. Chromium.

History.—In 1762 J. G. Lehmann, in a letter to the naturalist G. L. L. de Buffon, described a new mineral from Siberia. We now know this mineral to be crocoisite, or lead chromate. Both L. N. Vauquelin and Macquart, in 1789, failed to recognize in the mineral a new element, and both reported lead, iron, alumina, and a large amount of oxygen. However, in 1797, L. N. Vauquelin re-examined the mineral and concluded that the lead must be combined with a peculiar acid which he considered to be the oxide of a new metal. This he called chromium—from the Greek *χρῶμα* (*chroma*), colour—because its compounds are all coloured. In 1798 L. N. Vauquelin detected the new element in spinel and in smaragdite, and F. Tassart found chromium in chrome iron ore in 1799.

Occurrence.—Metallic chromium does not occur free in nature. It

occurs combined with oxygen in *chrome ochre*, which is chromium sesquioxide, Cr_2O_3 , associated with more or less earthy matters. Chromite, $\text{Fe}(\text{CrO}_2)_2$, is the chief ore of chromium. It also occurs as lead chromate in crocoite or *crocoisite*, PbCrO_4 . Traces occur in many minerals—emerald, jade, serpentine, etc.

Preparation of the metal.—Chromium metal can be prepared by reducing chromium sesquioxide with carbon in the electric furnace; or better, by the aluminothermic process, which is also called, after its inventor, the H. Goldschmidt's process (1905). An intimate mixture of chromium sesquioxide and aluminium powder, *A*, Fig. 187, is placed in a refractory clay crucible so that about two-thirds of the crucible is filled. A mixture of sodium or barium peroxide and aluminium powder is placed over this, as at *B*, Fig. 187. A piece of magnesium ribbon, *C*, is stuck into the latter mixture, and a layer of powdered fluorspar, *D*, is placed over all. The crucible is then set in a tray of sand and the magnesium ribbon, *C*, ignited. When the flame reaches the peroxide mixture, *B*, the aluminium is oxidized with explosive violence, and care must be taken to protect

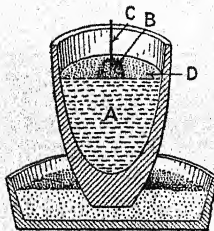


FIG. 187.—Alumino-Thermic Process for reduction of Oxides.

the face and hands accordingly. The heat of the combustion of the aluminium in the ignition mixture, *B*, starts the reaction between the chromic oxide and the aluminium, and it spreads throughout the entire mass in a few seconds. The chromic oxide is reduced to metal, and the aluminium is oxidized to alumina: $\text{Cr}_2\text{O}_3 + 2\text{Al} = 2\text{Cr} + \text{Al}_2\text{O}_3$. When the crucible is cold, a button of metallic chromium will be found on the bottom. The slag is nothing but fused alumina which has crystallized so as to form a kind of artificial corundum. This has been called *corubin* to distinguish it from natural corundum. The corundum slag is used as a

refractory and abrasive agent. When chromic oxide is reduced, the slag contains small artificial rubies. In Goldschmidt's works at Essen, about 100 kilograms of chromium are produced at a single charge. The process occupies less than half an hour. Manganese, iron, and many other metals can be produced in a similar manner. Titanium, alloyed with iron—ferro-titanium—is produced by the same process. So much heat is evolved during the reaction that even the most refractory metals and minerals are melted. Indeed, a temperature equivalent to that of the electric arc-furnace, 3000° , can be generated in a few minutes. The method is in some ways superior to the electric furnace process—it is quicker, cheaper, and the products are free from carbon contamination.

What amount of heat is generated when dry ferric oxide is reduced by aluminium, given that the heat of oxidation of iron to Fe_2O_3 is 195.6 Cals., and of aluminium to Al_2O_3 , 392.6 Cals.?

Heat developed by the oxidation of aluminium 392.6 Cals.

Heat absorbed by the reduction of ferric oxide 195.6 Cals.

Heat of the reaction $\text{Fe}_2\text{O}_3 + 2\text{Al} = \text{Al}_2\text{O}_3 + \text{Fe}$ 197.0 Cals.

Given the specific heat of the slags, etc., and the quantities of products, an approximate estimate of the temperature could be made. The specific heats here assumed to be constant are not so.

Properties of metallic chromium.—Chromium is a hard metal of a steel-grey colour. A 98–99 per cent. sample melted at 1520° , and boiled about 2200° . The pure metal probably melts at a much higher temperature—possibly 2000° . It might be added that chromium and many other rarer metals have not been prepared in a state approaching purity in quantities sufficiently large to determine their physical properties, and consequently, the physical properties of these elements reported in this and other text-books are to be understood “subject to revision.” The metal is fairly stable in air but oxidizes when heated to a high temperature, forming chromium sesquioxide, Cr_2O_3 . The metal dissolves in dilute hydrochloric and sulphuric acids, forming respectively chromous chloride and chromous sulphate, with the evolution of hydrogen. When placed in contact with nitric acid, the metal becomes inert or passive, for it is then no longer attacked by acids which dissolve it under normal conditions. The phenomenon of “passivity” is discussed in connection with iron.

Atomic weight.—The combining weight of chromium has been determined from the amount of chromium in silver and barium chromates; in chromium sulphate; in ammonium chromium alum; chromium chloride, potassium and ammonium dichromates, etc. The results show that if oxygen be 16, the combining weight of chromium lies somewhere between 51.6 and 53.5. The atomic weight is generally taken to be 52.1, $\text{O}=16$. This number agrees with Dulong and Petit’s rule, for the specific heat of chromium is 0.12, and $6.4 \div 0.12$ furnishes the number 53.3. This number is sufficiently close to 52.1 to show that 52.1 represents the atomic weight of chromium if Dulong and Petit’s rule applies to chromium.

Uses.—Ferro-chromium alloys are made containing over 60 per cent. of chromium, and less than 2 per cent. of carbon by smelting high-grade chromites in the electric furnace. Ferro-chromium is used in the manufacture of chrome-steel. Chrome-steel is a hard, tough, and dense metal with a high tensile strength. Steel with 1 to $1\frac{1}{2}$ per cent. of carbon and $2\frac{1}{2}$ to 4 per cent. of chromium is so hard that it cannot be worked by ordinary hardened tool steels—for example, it is drill proof. It can, however, be welded to iron, and rolled into sheets, etc. It is used in the manufacture of burglar proof safes, cutlery, stamp-mill shoes, crusher jaws, knuckles for car couplings, etc. It is superior to every known metal for the wearing parts of crushing and pulverizing mills. Alloys of chromium, nickel, and iron are used for the armour plates of war-ships. The alloy is hard and elastic, and even if a projectile does penetrate the armour plating the metal does not crack.

Chromite is used in making the hearths of steel furnaces since it can be used as a neutral refractory material between the basic (magnesian) bricks in the interior of the furnace, and the acidic (siliceous) bricks outside. Chromite bricks are not injured by contact with basic, nor with acid bricks; whereas acidic and basic bricks, when heated in contact with one another, are likely to fuse at the surfaces of contact owing to the formation of fusible silicates. The bricks are very refractory—softening between 2000° and 2100° —and do not crack by sudden heating and cooling.

Potassium and sodium chromates are used in dyeing; in the manufacture of pigments (chrome yellow, chrome red, Guignet’s green, etc.); in tanning leather, etc. If gelatine or glue be treated with a trivalent

chromium salt, it is rendered insoluble in hot water. Use is made of this in tanning leather from skins of animals. The chromates and dichromates do not do this, but if a mixture of gelatine or glue and a dichromate be exposed to light, the dichromate is slowly reduced, and the product "tans" the glue. Chromic trioxide is used in place of nitric acid in some voltaic batteries, etc.

§ 6. Molybdenum, Tungsten, and Uranium.

	Molybdenum, Mo.	Tungsten, W.	Uranium, U.
Atomic weight	96	184	238.5

Molybdenum, Mo.

The term *μόλυβδος* (molybdos) was applied by the Greeks to galena and other lead ores. Up to the middle of the 18th century, the mineral molybdate or molybdenite was supposed to be identical with graphite, then known as "plumbago" or "black lead." In 1778 C. W. Scheele, in his *Treatise on Molybdena*, showed that, unlike plumbago or graphite, molybdenite forms a "peculiar white earth" when treated with nitric acid. This he proved to have acid properties, and he called it "acidum molybdenæ," that is, molybdic acid; and he correctly considered the mineral molybdenite to be a molybdenum sulphide. In 1790 P. J. Hjelm isolated the element as a metallic powder by heating molybdic acid with charcoal. Molybdenum metal remains unchanged on exposure to air, but it oxidizes slowly at a red heat.

Molybdic trioxide, MoO_3 , is the most important compound of molybdenum. Like the analogous chromic trioxide, it behaves as an acid anhydride, forming molybdic acid, H_2MoO_4 , and salts are called molybdates. The mineral molybdenite—molybdenum sulphide—is first concentrated by separating the extraneous rocky matters by mechanical processes; the concentrate is roasted to convert it into molybdic oxide, MoO_3 , which can then be separated from the other oxides by digestion in aqueous ammonia. Soluble ammonium molybdate is formed. A solution of ammonium molybdate dissolved in an excess of nitric acid is used as a test for phosphates because it gives a yellow precipitate of ammonium phosphomolybdate with solutions containing phosphates. A similar precipitate is produced with arsenates. The composition of the precipitate varies a little with the conditions under which it is formed, so that the amount of As_2O_5 or P_2O_5 associated with a given amount of the precipitate is not always the same; consequently, in quantitative work, the precipitate is usually redissolved and the phosphorus re-precipitated as magnesium ammonium phosphate (*q.v.*) which is more easily controlled.

The variable composition of precipitated ammonium phosphomolybdate may seem to violate the "constancy of composition" test for distinguishing compounds from mixtures. It has been stated that "there are no exceptions to the law of constant composition," the real meaning of the phrase is that a compound is defined as a substance of constant composition, and we refuse to call any other substance a chemical compound. As P. Duhem (1902) has pointed out, so long as we do this there can be no exceptions to the law. In the present case, the difficulty is usually referred to the tendency of chromic, molybdic,

and tungstic acids to condense and form complex salts of the type $R'_2O.nCrO_3$; $R'_2O.nMoO_3$; and $R'_2O.nWO_3$, where n represents the numerical ratio between the CrO_3 , etc., and the R'_2O groups. Salts of acids more hydrated than H_2MoO_4 are known. Ordinary ammonium molybdate, for instance, is $(NH_4)_6Mo_7O_{24}.4H_2O$; ordinary sodium tungstate is $Na_{10}W_{12}O_{41}.28H_2O$, that is, $5Na_2O.12WO_3.28H_2O$. In addition to this, molybdic and tungstic acids can unite with one or more molecules of phosphoric, arsenic, silicic and other acids, forming still more complex acids; thus, we have phospho-, arseno-, arseni-, vanadi-, and silici-molybdic and tungstic acids, etc. Chromium, as we have seen, forms a similar series of acids; uranium exhibits but a slight tendency to form complex acids.

The more important oxides of molybdenum are: MoO , Mo_2O_3 , MoO_2 , MoO_3 , and in addition, several complex oxides appear to exist— Mo_2O_5 or $MoO_2.MoO_3$; Mo_3O_8 or $MoO_2.2MoO_3$; Mo_5O_{12} or $3MoO_2.2MoO_3$. The addition of nitric or hydrochloric acid to a concentrated solution of a molybdate precipitates the trioxide. The molybdates are reduced by zinc in acid solution to one of the lower oxides approximately, Mo_2O_3 , and at the same time the colour of the solution changes through various shades of violet, blue, and black. A reddish-brown molybdenum sulphide is precipitated by hydrogen sulphide in acid solution. The chlorides of molybdenum: $MoCl_2$, $MoCl_3$, $MoCl_4$, $MoCl_5$ and $MoCl_6$, are known. Hence molybdenum is 2-, 3-, 4-, 5-, and 6-valent.

Molybdenum steel is hard and less brittle than tungsten steel. Molybdenum is said to be far more effective than tungsten in the manufacture of hard steels. Molybdenum steel is used for making rifle barrels, propeller shafts, etc., and particularly high-speed tool steels. These steels, unlike ordinary carbon steels, have the peculiar property of retaining their "temper" when heated to a high degree, so that it is possible to make heavy cuts at high speed, for the steel can be heated to dull redness without impairing its quality. Molybdenum steel contains up to 10 per cent. of molybdenum. Some of the molybdenum and tungsten steels resist the action of acids unusually well, so that these *acid-proof steels* are useful in many chemical industries. Thus, an alloy containing about 60 per cent. of chromium, 35 per cent. of iron, and 2-3 per cent. of molybdenum is scarcely affected by dilute hydrochloric, nitric, or sulphuric acid, or by boiling aqua regia. Molybdenum compounds are also used as a blue pigment in porcelain painting; in silk and woollen dyeing; and in colouring leather and rubber—it has been, indeed, proposed for dyeing fabrics as a substitute for indigo. Ammonium molybdate is largely used in the determination of phosphorus in iron and steel laboratories.

Tungsten, W.

Up to the middle of the 18th century, the mineral scheelite—formerly called "tungsten" (heavy stone)—and wolframite were supposed to be ores of tin, but, in 1781, K. W. Scheele demonstrated that scheelite contains a peculiar acid, which he called tungstic acid, united with lime as a base. The same year, T. Bergmann recognized tungstic acid as an oxide of a new element, tungsten, which was isolated by J. J. y Don Fausto d'Elhuyar in 1783. When wolframite is mixed with calcium and ignited,

a vigorous reaction sets in, and metallic tungsten is obtained as a kind of regulus. When wolframite or scheelite is digested with concentrated hydrochloric acid, most of the iron and lime are removed, and tungsten trioxide, WO_3 , remains as a yellow insoluble powder. The metal is obtained by reducing the oxide or chloride in a current of hydrogen gas. Tungsten metal has been prepared in a highly ductile condition. It is very resistant to chemical agents—mineral acids and alkalis; its melting point is thought to be higher than that of any other metal; its tensile strength is greater than iron or nickel; it is non-magnetic; it can be drawn into as thin wires as any other metal; and its specific gravity 19.3 to 21.4 is 70 per cent. greater than that of lead. These valuable properties show that the metal promises to play an important part in the future. The metal is soluble in a mixture of hydrofluoric and nitric acids, in fused alkali nitrates, nitrites, and alkali peroxides. The tungsten compounds are somewhat similar to those of molybdenum. The chlorides— WCl_2 , WCl_4 , WCl_5 , and WCl_6 —show that tungsten may be 2-, 4-, 5-, and 6-valent. Fabrics which have been soaked in a solution of sodium tungstate and then dried do not burn with a flame but smoulder away slowly, hence sodium tungstate is used in making articles of clothing—e.g. flannelette—which the makers style “non-inflammable”—an insoluble tungstate is precipitated in the fibres of the fabric. Sodium tungstate is used as a mordant in dyeing. Lead tungstate has been used as a substitute for white lead in painting. The trioxide is used as a canary yellow pigment. Tungsten is also used in making high-speed steels—see “Molybdenum.” Tungsten filaments are also used in place of carbon for incandescent electric lamps. An old carbon filament lamp gave light of 16 candle power with a consumption of about 60 watts, a tungsten filament lamp will give 20 candle power with a consumption of 25 watts. Hence tungsten converts electric energy into light energy more efficiently than carbon, for there is less electric energy dissipated as heat during the conversion. Partially reduced tungstates furnish *bronzing powders—tungsten bronze*—for making paints to imitate bronzes.

Uranium, U.

The mineral pitchblende was formerly supposed to be an ore of zinc iron, or tungsten, but M. H. Klaproth (1789) proved that it contained what he styled a “half metallic substance” different from the three elements just named. This element was named “uranium” in honour of Herschel’s discovery of the planet Uranus in 1781. E. M. Péligot proved that Klaproth’s element was really an oxide of uranium, and he isolated the metal itself in 1842. When powdered pitchblende is digested with concentrated sulphuric acid, the uranium is converted into a sulphate soluble in water. The metal is obtained by reducing the chloride with sodium. Many of the more refractory metals—titanium, vanadium, etc.—are made by reducing their chlorides with hydrogen or sodium. An inconveniently high temperature is needed when hydrogen is used, and it is difficult to keep sodium free from oxidation. The chloride is therefore reduced by heating it with sodium in a steel-bomb screwed down tight. M. Billy (1914) found that when sodium hydride is used as a reducing agent, the reduction occurs at a comparatively low temperature—e.g. in a glass

vessel heated to, about 800° . The resulting metal is said to have a high degree of purity.

Uranium forms five oxides— U_2O_3 , UO_2 , U_3O_8 , UO_3 , UO_4 . The chlorides UCl_3 , UCl_4 , UCl_5 , UO_2Cl_2 , and a fluoride UF_6 are known. Hence uranium is 3-, 4-, 5-, and 6-valent. Uranium is quadrivalent in the uranous salts and sexivalent in the uranic salts. The uranic salts are derived from uranic acid, $UO_2(OH)_2$. Thus, uranic nitrate is $UO_2(NO_3)_2 \cdot 6H_2O$; uranic chloride is UO_2Cl_2 , etc. The group UO_2 is generally called uranyl, and the salts just named are respectively uranyl nitrate and uranyl chloride. Uranic trioxide, UO_3 , has also acidic properties, for it forms salts mono-, di-, tri-, and tetra-uranates, analogous in constitution with the corresponding chromates (p. 561). Thus sodium diuranate is $Na_2U_2O_7 \cdot 6H_2O$, etc. Sodium diuranate is called uranium yellow, and is used, as well as uranium oxide, for colouring glass and pottery glazes. Yellow uranium glass and some uranium salts have a yellowish green fluorescence. Uranium compounds are radioactive. Uranium is also used in the manufacture of incandescent mantles. When alloyed with steel, it makes the metal tough and hard. Solutions of uranyl nitrate or uranyl acetate are used in the volumetric determination of phosphoric acid. If a solution containing a known amount of, say, uranyl nitrate be added to a solution containing soluble phosphates, a greenish yellow precipitate of uranyl ammonium phosphate, $UO_2(NH_4)PO_4$, insoluble in acetic acid, will be formed as long as any phosphate remains in the solution. When the phosphates are all precipitated, any further addition of the standard solution of uranyl nitrate will cause the solution to give a brown coloration when a drop is brought in contact with a drop of a solution of potassium ferrocyanide on, say, a white plate.

Molybdenum is distributed fairly widely over the earth's crust. It occurs as *molybdenite*, that is, molybdenum sulphide, MoS_2 , *wulfenite*, lead molybdate, $PbMoO_4$; is a hard, brittle, white mineral; and *molybdcite*, or *molybdic ochre*, is a straw-yellow soft mineral supposed to be a hydrated iron molybdate. Uranium occurs as *pitchblende*, or *uraninite*, U_3O_8 ; as *carnotite*, $K_2O \cdot 2U_2O_3 \cdot V_2O_5 \cdot 3H_2O$. Tungsten occurs as *wolfram*, $FeWO_4$, and *scheelite*, $CaWO_4$. The metals are made by reducing the oxides or salts with reducing agents—carbon, potassium cyanide, hydrogen, sodium, or aluminium.

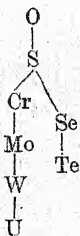
§ 7. The Relationship of the Chromium Family.

The metals chromium, molybdenum, tungsten, and uranium are difficult to fuse, their physical properties are indicated in the following table:—

TABLE XXXII.—PHYSICAL PROPERTIES OF THE CHROMIUM FAMILY.

	Chromium.	Molybdenum.	Tungsten.	Uranium.
Atomic weight . . .	52.1	98	184	238.5
Specific gravity . . .	6.74	9.01	19.13	18.7
Atomic volume . . .	7.7	10.6	9.6	12.7
Melting point . . .	1510°	c. 2500°	3000°	1500° (?)

The metals combine directly with oxygen, sulphur, nitrogen, and the halogens when heated in an atmosphere of these gases. The oxides are difficult to reduce. The trioxides— CrO_3 , MoO_3 , WO_3 , UO_3 —



are acidic, and they form salts of the type K_2RO_4 as indicated above. The formation of complex acids is also noteworthy. The relationship of the chromium family with the sulphur family is brought out by the analogy between potassium chromate and sulphate; sodium chromate and sulphate; sulphuryl and chromyl chlorides; dichromates, disulphates, and diuranates, etc. Thus, potassium sulphate, selenate, chromate, and manganate are isomorphous; and a series of tungstates, molybdates, selenates, and chromates of the type $3\text{Na}_2\text{RO}_4 \cdot \text{Li}_2\text{RO}_4 \cdot 12\text{H}_2\text{O}$ —where R represents sexivalent S, Se, Cr, Mo, or W—crystallize in the hexagonal system.

Starting with the elements with the smallest atomic weight, the relationship between the chromium and sulphur families is generally shown by a scheme like that in the margin. The fault with this scheme is that it makes the relationship appear closer than the known facts would lead us to suppose, and the student might reasonably think that there is a tendency to go a-begging for far-fetched analogies between unrelated elements. Chromium also has some analogies with metals of other groups, e.g. aluminium and trivalent iron (*q.v.*).

Questions.

1. The oxide of chromium, Cr_2O_3 , and the hydroxide, $\text{Cr}(\text{OH})_3$, are weak basic and still more weakly acidic. How does this statement accord with the fact that the salt, $\text{Cr}_2(\text{SO}_4)_3$, is stable in water while the corresponding carbonate and sulphide are not stable? What compound is formed when sodium carbonate is added to a solution of a chromic salt?—*Massachusetts Inst. Technology, U.S.A.*

2. What reaction takes place when thermit is ignited? What uses are made of thermit? How is metallic aluminium prepared commercially?—*Cornell Univ., U.S.A.*

3. Mention any conditions which you know to affect the colour of metallic salts in solution. Discuss the colour phenomena of solutions of cupric, chromic, and ferric salts.—*St. Andrews Univ.*

4. What happens when sodium hydroxide is mixed with a solution of a chromic salt? Of a potassium chromate solution? Of a potassium dichromate solution? Give four reactions which distinguish chromium from all other elements.—*Amherst Coll., U.S.A.*

5. Describe how you would prepare (a) barium chloride crystals from heavy spar; (b) potassium dichromate from chromium sesquioxide; (c) sodium carbonate and sodium bicarbonate from caustic soda.—*Aberystwyth Univ.*

6. Nilson and Petterson found that 0.0859 gram of chromium sesquichloride when volatilized at 1200°C . displaced 12.049 c.c. of gas (reduced to 0°C . and 760 mm.), calculate the density of the vapour compared with air and the formula to which it corresponds. How would you prepare potassium chromate, potassium bichromate, and chromic acid?—*Owens Coll.*

7. On heating an unknown quantity of potassium dichromate with a concentrated solution of hydrogen chloride, and passing the evolved gas into a solution of potassium iodide, 1.25 grams of iodine were liberated: what was the weight of the dichromate used? ($\text{K}=39$, $\text{I}=127$, $\text{Cr}=52$.)

8. What use has been made of aluminium for the isolation of other metals? The heat of combustion of aluminium per gram is 7250, of sodium 2170, and of carbon 8000 calories. From these data determine whether it is possible to reduce aluminium or sodium from their oxides by carbon.—*Board of Educ.*

9. Calculate the number of cubic centimetres of decinormal potassium dichromate solution required to oxidize 0.5 gram of iron dissolved in dilute hydrochloric acid. In an actual experiment 88.5 cubic centimetres were used. How do you account for the difference? ($K=39$, $Cr=52$, $O=16$, $Fe=56$.)—*Trinity Coll.*

10. Sulphuric acid or alcohol can be dissolved in water in all proportions, and it is accordingly stated that "these must be considered as conditions of mixture rather than of chemical combination." Discuss this statement.

11. Contrast the properties of oxygen and ozone, of the two oxides of sulphur, and of chromium sesquioxide and chromium trioxide, and point out how much comparisons justify the belief that the properties of substances are very largely dependent on composition as well as on the nature of the elements they contain.—*London Univ.*

12. Sometimes anhydrous salts have not the same colour as their dilute solutions. Give an illustration with explanation.—*Cape Univ.*

CHAPTER XXV

MANGANESE

§ 1. Manganese Oxides.

Manganese dioxide, MnO_2 .—The mineral *pyrolusite* commonly contains from 70 to 90 per cent. of manganese dioxide, MnO_2 , contaminated with more or less iron, alumina, silica, lime, baryta, and maybe cobalt. When heated, manganese dioxide loses oxygen and changes to manganese sesquioxide, Mn_2O_3 , and then to mangano-manganic oxide, Mn_3O_4 . Cold concentrated hydrochloric acid gives a dark brown liquid, and very little chlorine is evolved. The cold solution probably contains manganese trichloride, MnCl_3 , and possibly also a little manganese tetrachloride, MnCl_4 ; but the composition of the liquid is not definitely known. In any case, chlorine gas (*q.v.*) is evolved when the liquid is warmed, and manganous chloride, MnCl_2 , remains in solution. When heated with sulphuric acid, a solution of manganese sulphate, MnSO_4 , is formed and oxygen gas is evolved: $2\text{MnO}_2 + 2\text{H}_2\text{SO}_4 \rightarrow 2\text{MnSO}_4 + 2\text{H}_2\text{O} + \text{O}_2$.

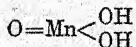
Manganese dioxide appears to be a feeble acidic oxide; but since it dissolves in acids with the evolution of oxygen or its equivalent, it is probably an extremely feeble basic oxide, if at all. There is no direct evidence of the existence of manganese tetrachloride, MnCl_4 , in aqueous solution, but double salts, $\text{MnCl}_4 \cdot 2\text{KCl}$, or rather complex salts, K_2MnCl_6 , are known, *cf.* p. 303. Manganese dioxide does not give hydrogen peroxide with acids, and it is not therefore a "superoxide," $\text{Mn} < \overset{\text{O}}{\text{O}}$; it is usually represented as a "polyoxide," $\text{O}=\text{Mn}=\text{O}$, although some consider it to be $(\text{MnO}_2)_2$ —a manganous manganate, $\text{Mn}=\text{MnO}_4$ —as shown graphically on p. 579. The constitution of manganese dioxide has not yet been made clear.

Manganic sulphate, $\text{Mn}_2(\text{SO}_4)_3$, is obtained as a dark green powder when manganese dioxide is gently heated with concentrated sulphuric acid. The manganic sulphate is immediately hydrolyzed by water, and it forms a series of "alums," isomorphous with aluminium alums, when treated with alkaline sulphates. Thus, manganic potassium alum, $\text{Mn}_2(\text{SO}_4)_3 \cdot \text{K}_2\text{SO}_4 \cdot 24\text{H}_2\text{O}$, crystallizes in violet octahedral crystals. The alums are rather more stable than manganic sulphate itself, but even then the potassium and ammonium manganic alums are so unstable that it is doubtful if they have been made pure, for they are hydrolyzed by water, forming manganic hydroxide, $\text{Mn}(\text{OH})_3$. Cesium and rubidium manganic alums are fairly stable.

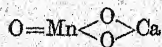
Hydrated manganese dioxide, $\text{MnO}_2 \cdot \text{H}_2\text{O}$, or manganous acid,

H_2MnO_3 .—When an alkaline hypochlorite or hypobromite (or an alkaline hydroxide with chlorine or bromine water) is added to an aqueous solution of a manganous salt, say, manganous chloride, MnCl_2 , the so-called “hydrated manganese oxide, $\text{MnO}_2 \cdot \text{H}_2\text{O}$,” is precipitated, just as lead dioxide is precipitated from a lead salt solution under similar conditions. The reaction is represented: $2\text{MnCl}_2 + 4\text{NaOCl} = 2\text{MnO}_2 + 4\text{NaCl} + 2\text{Cl}_2$. The compound $\text{MnO}_2 \cdot \text{H}_2\text{O}$ is possibly a manganous acid, H_2MnO_3 . A colloidal solution of this acid is easily obtained. Many compounds of manganese dioxide with the basic oxides are known. The “manganese mud” formed by blowing air through a solution containing a mixture of lime with a solution of manganese chloride is supposed to contain calcium manganite— $\text{CaO} \cdot \text{MnO}_2$; or CaMnO_3 . Calcium manganite is insoluble in water and slowly settles as a black mud—p. 277.

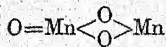
Manganese sesquioxide, Mn_2O_3 .—Also called red oxide of manganese. It occurs in nature as *brownite*, Mn_2O_3 , or rather $3\text{Mn}_2\text{O}_3 \cdot \text{MnSiO}_3$. Manganese sesquioxide is obtained as a black powder when any other oxide of manganese is heated to about 900° in a current of oxygen. The corresponding manganic hydroxide, $\text{MnO} \cdot \text{OH}$, is formed when manganic sulphate is decomposed by water. This compound is converted into manganese dioxide when calcined to about 300° . The salts, manganic sulphate and manganic chloride, have been described. With concentrated hydrochloric acid, the hydroxide seems to furnish manganese trichloride (*q.v.*). This shows that manganese sesquioxide possesses basic functions. Neither the oxide nor the hydroxide seems to dissolve in cold sulphuric acid; hot dilute sulphuric acid forms manganous sulphate, MnSO_4 , and leaves manganese dioxide insoluble. With hot nitric acid, manganous nitrate and manganese dioxide are formed: $2\text{MnO} \cdot \text{OH} + 2\text{HNO}_3 \rightarrow \text{Mn}(\text{NO}_3)_2 + \text{MnO}_2 + 2\text{H}_2\text{O}$. Hence, this oxide is sometimes represented as $\text{MnO} \cdot \text{MnO}_2$, or better as manganous manganite, $\text{Mn} \cdot \text{MnO}_3$, analogous with calcium manganite. The manganites are thus represented graphically:



Manganous acid.



Calcium manganite.

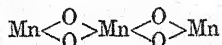


Manganous manganite.

This is hypothesis; however, trivalent manganese compounds are known, and manganese sesquioxide may be one of them: $\text{O}=\text{Mn}-\text{O}-\text{Mn}=\text{O}$.

Mangano-manganic oxide, Mn_3O_4 .—This oxide occurs in nature in red prismatic crystals of *hausmannite*, and it is formed as a brownish-red powder when any of the manganese oxides are ignited in air. It can be obtained in a crystalline condition by heating the powdered oxide in a current of hydrogen chloride. Mangano-manganic oxide is not a simple basic oxide, for its behaviour towards acids leads to the formation of a soluble manganous salt, and insoluble manganese dioxide: $\text{Mn}_3\text{O}_4 + 2\text{H}_2\text{SO}_4 \rightarrow 2\text{MnSO}_4 + \text{H}_4\text{MnO}_4$. The hydrate then decomposes into manganese dioxide and water. Hence mangano-manganic oxide appears to be a compound oxide— $2\text{MnO} \cdot \text{MnO}_2$, or $\text{MnO} \cdot \text{Mn}_2\text{O}_3$; or better still, manganese orthomanganite: Mn_2MnO_4 . The action of nitric acid indicated above may then be written: $\text{Mn}_2\text{MnO}_4 + 4\text{HNO}_3 \rightarrow 2\text{Mn}(\text{NO}_3)_2 + \text{H}_4\text{MnO}_4$. The oxide, Pb_3O_4 (*q.v.*), is generally supposed to have the constitution, Pb_2PbO_4 , or $2\text{PbO} \cdot \text{PbO}_2$; and Fe_3O_4 is supposed to be $\text{FeO} \cdot \text{Fe}_2\text{O}_3$, since no oxide corresponding with quadrivalent iron— FeO_2 —is known to exist. Hence our theory of the constitution of Mn_3O_4 depends

upon what view is taken of the constitution of Mn_2O_3 . A red solution, containing manganous and manganic sulphates, is formed when Mn_3O_4 is treated with concentrated sulphuric acid, this corresponds with the formula $\text{MnO} \cdot \text{Mn}_2\text{O}_3$ for mangano-manganic oxide; and the formation of manganese dioxide and manganous salt when mangano-manganic oxide is treated with dilute sulphuric or nitric acid agrees with Mn_2MnO_4 , or $2\text{MnO} \cdot \text{MnO}_2$, that is, with



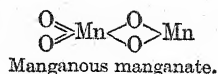
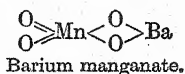
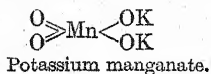
Manganous oxide, MnO .—When manganous chloride is treated with an alkaline hydroxide, in the absence of air, a colourless flocculent precipitate of manganous hydroxide, $\text{Mn}(\text{OH})_2$, is formed. This quickly oxidizes probably to green manganic hydroxide, $\text{Mn}(\text{OH})_3$. The hydroxide is slightly soluble in water. If ammonium salts be present, a soluble complex salt, say, $(\text{NH}_4)_2\text{MnCl}_4$, is formed. Hence, the complete precipitation of manganese hydroxide by ammonia is prevented by the simultaneous formation of ammonium salts. If the ammoniacal solution of manganese hydroxide be exposed to air, manganic hydroxide, $\text{Mn}(\text{OH})_3$, is slowly precipitated.

If sodium carbonate be added to a solution of a manganous salt, manganous carbonate, MnCO_3 , is precipitated. If manganous carbonate or hydroxide be heated in the absence of air, or if any of the hydroxides of manganese be heated in a current of hydrogen gas, Mn_2O_3 is first formed; at 280° Mn_3O_4 is produced; and finally a greenish powder of manganous oxide, MnO , is obtained. This oxidizes rapidly on exposure to the air. The manganous salts are readily obtained by treating the carbonate or hydroxide with the proper acid. The manganous salts are pink in colour, the aqueous solutions are almost colourless. Unlike the ferrous salts, the manganous salts are stable in the solid condition, and also in neutral or acid solutions. Manganous chloride and sulphate both form double or complex salts with the alkali salts— $\text{MnCl}_2 \cdot 2\text{NH}_4\text{Cl} \cdot \text{H}_2\text{O}$; $\text{MnSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$. Manganese and iron form similar isomorphous double salts, p. 304.

§ 2. Manganates and Permanganates.

Manganates.—When manganese dioxide is fused with potassium or sodium hydroxide, in the absence of air, a part of the manganese is converted into the corresponding manganate; e.g. $3\text{MnO}_2 + 2\text{KOH} = \text{K}_2\text{MnO}_4 + \text{Mn}_2\text{O}_3 + \text{H}_2\text{O}$. If air be present, and if some oxidizing agent—potassium nitrate or chlorate—be associated with the mixture, the manganese dioxide can be nearly all converted into potassium manganate: $2\text{MnO}_2 + 4\text{KOH} + \text{O}_2 = 2\text{K}_2\text{MnO}_4 + 2\text{H}_2\text{O}$. The fused mass has a dark green colour, and when diluted with a small quantity of water, it furnishes a dark green solution from which dark green crystals of potassium manganate, K_2MnO_4 , can be obtained by allowing the solution to evaporate, at ordinary temperatures, in vacuo. Sodium manganate can be made by a similar process using sodium in place of potassium compounds; and it can also be made by fusing manganese dioxide and other manganese compounds with sodium peroxide. The sodium manganate, $\text{Na}_2\text{MnO}_4 \cdot 10\text{H}_2\text{O}$, has a composition resembling Glauber's salt, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$. The

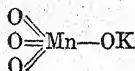
isomorphism of sodium manganate with sodium sulphate and sodium chromate makes it probable that all these compounds have a similar constitution, and that like sulphur in the sulphates, manganese is sexivalent in the manganates. By analogy with the sulphates, therefore, the graphic formula of potassium manganate is written :



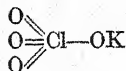
Permanganates.—When the green concentrated solution of potassium manganate is gently warmed, or largely diluted with water, the green colour changes to pink owing to the formation of a solution of potassium permanganate, KMnO_4 , and a precipitation of hydrated manganese dioxide, $\text{MnO}_2 \cdot \text{H}_2\text{O}$, thus: $3\text{K}_2\text{MnO}_4 + 3\text{H}_2\text{O} = 2\text{KMnO}_4 + \text{MnO}_2 \cdot \text{H}_2\text{O} + 4\text{KOH}$. It is supposed that the potassium manganate is first hydrolyzed, forming manganic acid: $\text{K}_2\text{MnO}_4 + 2\text{H}_2\text{O} = 2\text{KOH} + \text{H}_2\text{MnO}_4$; and that the manganic acid is so unstable that it is decomposed at once: $3\text{H}_2\text{MnO}_4 = 2\text{HMnO}_4 + \text{MnO}_2 \cdot \text{H}_2\text{O} + \text{H}_2\text{O}$. The manganate is thus self-oxidized and self-reduced, for the manganate is decomposed into a compound richer in oxygen and at the same time into a compound poorer in oxygen. One part of the compound is oxidized at the expense of the oxygen in another part. Several examples of this phenomenon have already been given—the decomposition of hypochlorites into chlorates and chlorides; the ignition of a chlorate gives a perchlorate and a chloride; while alkaline thiosulphates and sulphites give, on calcination in the absence of air, sulphates and sulphides. We shall also find later that nitrous acid furnishes nitric acid and nitric oxide; hypophosphorous and phosphorous acids give phosphoric acid and phosphine. Potassium manganate is not hydrolyzed in alkaline solutions, and it is supposed that pure water will not hydrolyze the solution. If a small trace of acid be present, even carbonic acid derived from the atmosphere, the hydrolysis takes place. Hence if carbon dioxide be passed through an aqueous solution of potassium manganate, the latter is converted into a permanganate: $\text{K}_2\text{MnO}_4 + \text{H}_2\text{CO}_3 = \text{K}_2\text{CO}_3 + \text{H}_2\text{MnO}_4$, and the manganic acid is then decomposed as indicated above. The oxidization of acidic solutions of manganese salts may be effected by persulphates, by lead peroxide (*Crum's test*), and by sodium bismuthate. These reactions are utilized for the colorimetric determination of manganese. In addition solutions of potassium manganate can be oxidized by means of chlorine, ozone, hydrogen peroxide, and other oxidizing agents. Potassium permanganate is made by passing carbon dioxide through the aqueous solution obtained by extracting a fused mixture of manganese dioxide and potassium hydroxide: $3\text{K}_2\text{MnO}_4 + 2\text{CO}_2 \rightarrow 2\text{K}_2\text{CO}_3 + 2\text{KMnO}_4 + \text{MnO}_2$. The separation of manganese dioxide is avoided by oxidizing the alkali manganate solution as just described, by electrolyzing it between iron or nickel electrodes: $2\text{K}_2\text{MnO}_4 + 2\text{H}_2\text{O} \rightarrow 2\text{KMnO}_4 + 2\text{KOH} + \text{H}_2$. The potassium permanganate is separated by crystallization, and the hydroxide recovered by evaporation and used for the next fusion.

Potassium permanganate forms dark purple, almost black crystals with a greenish lustre. The crystals are isomorphous with potassium perchlorate, KClO_4 . The isomorphism makes it probable that manganese

in potassium permanganate, like chlorine in potassium perchlorate, is septivalent, and that the compounds have the graphic formulæ:



Potassium permanganate.

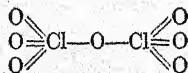
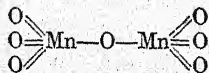


Potassium perchlorate.

The crystals are moderately soluble in water: 100 grams of water at 0° dissolve 2.83 grams of the salt; and at 50°, 16.89 grams. The saturated solution is an intense purple colour.

When heated to about 240°, potassium permanganate furnishes oxygen and potassium manganate: $6\text{KMnO}_4 = 3\text{K}_2\text{MnO}_4 + 3\text{MnO}_2 + 3\text{O}_2$. With concentrated sulphuric acid, it decomposes explosively: $4\text{KMnO}_4 + 6\text{H}_2\text{SO}_4 = 2\text{K}_2\text{SO}_4 + 4\text{MnSO}_4 + 6\text{H}_2\text{O} + 5\text{O}_2$; but it is not decomposed by very dilute sulphuric acid, hot or cold. If a hot aqueous solution of silver nitrate and potassium permanganate be cooled, silver permanganate, AgMnO_4 , separates; and if the solution of silver permanganate be treated with barium chloride, barium permanganate, $\text{Ba}(\text{MnO}_4)_2$, and insoluble silver chloride are formed.

Permanganic acid and its anhydride.—Potassium permanganate is a salt of permanganic acid, HMnO_4 . Permanganic acid is best made by adding just sufficient sulphuric acid to barium permanganate to convert all the barium into barium sulphate. On evaporating the filtered solution, violet crystals of permanganic acid are obtained. Like perchloric acid, HClO_4 , permanganic acid, HMnO_4 , is a powerful oxidizing agent, and like perchloric acid, it decomposes in contact with organic matter. If solid potassium permanganate be cautiously added to well-cooled concentrated sulphuric acid, a green oily liquid is obtained which appears to be $(\text{MnO}_3)_2\text{SO}_4$. If a little water be added to this solution, well cooled, a dark reddish-brown liquid separates which does not solidify at -20° . This is supposed to be manganese heptoxide, Mn_2O_7 : $(\text{MnO}_3)_2\text{SO}_4 + \text{H}_2\text{O} = \text{Mn}_2\text{O}_7 + \text{H}_2\text{SO}_4$, or graphically, by analogy with chlorine heptoxide:

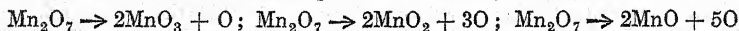


Manganese heptoxide is very unstable, and decomposes with violence when warmed, forming a lower oxide and oxygen: $2\text{Mn}_2\text{O}_7 = 4\text{MnO}_2 + 3\text{O}_2$. A mixture of sulphur, or phosphorus with potassium permanganate, is violently explosive. Manganese heptoxide is permanganic anhydride, because, when treated with cold water and sulphuric acid, it regenerates the green solution of $(\text{MnO}_3)_2\text{SO}_4$. Perhaps the analogy between the permanganates and the perchlorates has been pushed too far because the manganates do not seem to be related to the chlorates. The parallelism between the manganates and permanganates is closer than between the chlorates and perchlorates.

Manganic acid and its anhydride.—If solid potassium permanganate be dissolved in concentrated sulphuric acid, and the green solution—probably of $(\text{MnO}_3)_2\text{SO}_4$ —be dropped upon dry sodium carbonate, violet fumes are evolved. These fumes may be condensed to a red viscid solid not yet proved to be MnO_3 , manganic anhydride or manganese trioxide.

Some consider the red solid to be a mixture of manganese heptoxide with a lower oxide, say MnO_2 . Those who consider it to be an impure trioxide say that the permanganic anhydride is probably decomposed by the reaction: $2\text{Mn}_2\text{O}_7 = 4\text{MnO}_3 + \text{O}_2$. Manganese trioxide is decomposed by water: $3\text{MnO}_3 + \text{H}_2\text{O} = 2\text{HMnO}_4 + \text{MnO}_2$. Manganic acid, H_2MnO_4 , might be expected to be a product of this reaction; but manganic acid has not been prepared in the pure state.

Oxidizing action of the permanganates.—Potassium permanganate is not acted upon by pure sodium or potassium hydroxides in aqueous solutions, but the commercial alkaline hydroxides generally contain sufficient impurities to reduce some of the permanganate. When heated with an alkali, potassium permanganate reverts to potassium manganate: $4\text{KMnO}_4 + 4\text{KOH} = 4\text{K}_2\text{MnO}_4 + 2\text{H}_2\text{O} + \text{O}_2$; if a reducing agent be present, the alkaline permanganate solution is further reduced to manganese dioxide: $4\text{KMnO}_4 + 2\text{H}_2\text{O} = 4\text{MnO}_2 + 4\text{KOH} + 3\text{O}_2$; while in acidified solutions, the reduction proceeds still further, and a manganous salt is formed: $4\text{KMnO}_4 + 6\text{H}_2\text{SO}_4 = 2\text{K}_2\text{SO}_4 + 4\text{MnSO}_4 + 6\text{H}_2\text{O} + 5\text{O}_2$. Hence 2 gram-molecules of potassium permanganate contain the equivalent of 80 grams of available oxygen. There are thus three stages in the reduction of potassium permanganate corresponding with the separation of 1, 3, and 5 atoms of oxygen per two molecules of salt; the decomposition products are respectively potassium manganate, manganese dioxide, and a manganous salt. In symbols:



The first takes place in alkaline solutions; the second in alkaline and reducing solutions; and the third in acid solutions.

Oxalic acid, ferrous sulphate, sulphurous acid, hydrogen peroxide, nitrous acid, etc., reduce potassium permanganate to a manganous salt in an acid solution. With hydrogen peroxide, oxygen is evolved: $5\text{H}_2\text{O}_2 + 2\text{KMnO}_4 + 3\text{H}_2\text{SO}_4 \rightarrow 2\text{MnSO}_4 + \text{K}_2\text{SO}_4 + 8\text{H}_2\text{O} + 5\text{O}_2$. The reaction with oxalic acid, $(\text{COOH})_2$, is symbolized: $2\text{KMnO}_4 + 5(\text{COOH})_2 + 3\text{H}_2\text{SO}_4 \rightarrow 2\text{MnSO}_4 + 10\text{CO}_2 + \text{K}_2\text{SO}_4 + 16\text{H}_2\text{O}$. The oxalic acid actually employed has 2 molecules of water of crystallization: $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$. The reaction with ferrous sulphate is represented: $2\text{KMnO}_4 + 10\text{FeSO}_4 + 8\text{H}_2\text{SO}_4 \rightarrow 2\text{MnSO}_4 + 5\text{Fe}_2(\text{SO}_4)_3 + \text{K}_2\text{SO}_4 + 8\text{H}_2\text{O}$. These reactions giving the terminal products are used in analytical chemistry. Since the solution of the manganous salt is almost colourless, if a solution of potassium permanganate containing a known amount of the salt per litre be added from a burette, the permanganate is decolorized as fast as it is added to the reducing agent, until all the latter has been oxidized. The appearance of a permanent pink coloration due to the permanganate shows that all the reducing agent is destroyed. The solution must be kept acid or a precipitate of hydrated manganese dioxide will be formed.

§ 3. Manganese.

History.—Manganese appears to have been used by the ancient Egyptians and Romans for bleaching glass, for their glass often contains the equivalent of up to 2 per cent. manganese oxide. Pliny mentions its use for this purpose under the name "magnes"; he considered it to be

a variety of loadstone, i.e. a variety of magnetic iron ore. For reasons stated in connection with magnesium, the term "magnesia" in old books may refer to manganese oxide. B. Valentine, and chemists generally towards the end of the 18th century, believed wad to be an ore of iron. J. H. Pott (1740) proved that pyrolusite proper does not contain iron, and furnishes a number of salts quite different from those obtained with the iron oxides. K. W. Scheele (1774) made an important investigation on manganese (*vide* chlorine), and T. Bergman (1774) suspected that some metal lay concealed in the mineral earth, partly because of its specific gravity, partly because of the peculiar colour it imparts to glass, and partly because solutions of the earth in acid give a precipitate on the addition of alkali carbonate, a circumstance, he said, which takes place with all metals but not with any earth. Bergman reduced the metal earth with charcoal and obtained a metal regulus; a purer sample was afterwards isolated by J. F. John in 1807.

Occurrence.—The metal manganese does not occur free in nature. Much of the manganese ore comes from Russia, and the highest grade picked pyrolusite runs 50 per cent. manganese. Some ore comes from several other countries. The chief minerals are the oxides *pyrolusite*, MnO_2 ; *braunite*, Mn_2O_3 ; *hausmannite*, Mn_3O_4 ; *manganite*, $\text{Mn}_2\text{O}_3 \cdot \text{H}_2\text{O}$. The carbonate, MnCO_3 , *rhodochrosite* or *rhodochrosite*, is often associated with siderite (FeCO_3); manganese also less frequently occurs as sulphide, *manganese blende*, MnS ; and silicate, *rhodonite*, MnSiO_3 . *Franklinite* usually contains manganese as well as iron, and, after the zinc has been extracted, is used in making spiegeleisen. *Wade* is an impure mixture of manganese oxides often found in damp low-lying places. *Wad* is supposed to be a decomposition product of the manganese minerals. Minute quantities of manganese occur in water, plants, and animals; and traces also appear to be the colouring agent of many amethyst-coloured minerals.

Preparation of the metal.—Metallic manganese has been obtained by reducing the oxide with carbon when a very high temperature is required. It is far better to mix dry manganese dioxide with dry aluminium powder in a crucible, and to ignite the mixture as in Goldschmidt's process, Fig. 187, for the reduction of chromium oxide: $3\text{MnO}_2 + 4\text{Al} = 2\text{Al}_2\text{O}_3 + 3\text{Mn}$. The metallic manganese and alumina are melted by the high temperature, and the metal collects at the bottom of the crucible.

Properties of the metal.—Manganese is a grey metal with a reddish tinge like bismuth. The metal is brittle and harder than iron. It has a density of 8.0; a sample of 97 to 98 per cent. purity melts at 1225° and boils at 1900° . The metal volatilizes in the electric arc furnaces. Manganese is superficially oxidized when exposed to the air, and it decomposes in contact with water with the evolution of hydrogen. When heated in nitrogen, it forms manganese nitride, Mn_3N_2 ; and if heated in ammonia it forms another nitride, Mn_2N_2 . It also combines directly with carbon to form manganese carbide, Mn_3C . It reacts with nitrogen at about 1200° . Manganese readily dissolves in acids—hydrochloric acid, nitric acid, and acetic acid—forming manganous salts with the evolution of hydrogen gas. The valency of manganese is noteworthy since it acts as a bi-, ter-, quadri-, sexi-, and septivalent element. Manganese also forms an unusual number of definite oxides, more indeed than any other element.

Manganous oxide, MnO	Basic, forms manganous salts.
Manganic oxide, Mn_2O_3	Neutral or mixed oxide; probably a salt.
Manganese sesquioxide, Mn_2O_3	Basic, forms manganic salts.
Manganese dioxide, MnO_2	Basic dioxide and acidic; forms manganites.
Manganese trioxide, MnO_3	Acidic, forms manganates.
Manganese heptoxide, Mn_2O_7	Acidic, forms permanganates

Atomic weight.—Analyses of manganese chloride and bromide, silver permanganate, roasting manganous sulphate to oxide, and reducing manganese sulphate to manganese sulphide in a stream of hydrogen sulphide, all show that the combining weight of manganese (oxygen = 16) lies between 54.925 and 55.014; the best representative value is taken to be 54.93. This agrees with the estimation of the atomic weight by Dulong and Petit's rule.

Relation of manganese to the halogens.—Manganese does not form a family group with other elements with similar characters, as is the case with the halogens. Manganese, however, is usually classed with the halogens, but there are not many common properties between them. The similarity between the halogens and manganese virtually begins and ends with compounds of the highest oxide, Mn_2O_7 , which itself is strikingly like Cl_2O_7 . The corresponding acids are both monobasic, powerful oxidizing agents, and form isomorphous salts. There is a very great contrast between the lower oxides of chlorine and manganese, and between the elements chlorine and manganese. The relationship between chromium, manganese, and iron is much closer. Thus the isomorphism of the manganites and chromites; the isomorphism between the manganic and ferric alums, etc. The metals chromium, manganese, and iron also have many similar properties.

Uses.—Manganese dioxide is used as an oxidizing agent; in the manufacture of chlorine and bromine. It is used in decolorizing glass stained a yellowish tinge by the traces of "ferric silicate" present, for the violet colour of manganese silicate masks the complementary yellow tint of the iron. Manganese dioxide is also used as a "drier" for paints and varnishes; as a depolarizer in battery cells; colouring pottery bodies and glazes; etc. Wad is used in the manufacture of paint. A crude mixture of sodium manganate and permanganate is made by fusing sodium hydroxide with pyrolusite, and sold as a disinfectant under the name "Condy's fluid." Its "disinfecting" qualities depend upon its oxidizing properties.

Manganese is used in the manufacture of manganese bronze (*q.v.*), and also in the manufacture of iron and manganese steel. Manganese alloyed with less than 20 per cent. of iron is called spiegeleisen and with between 20 and 90 per cent. of iron, ferromanganese which can be made by reducing the oxides with carbon in an ordinary blast furnace, or in an electric furnace. Ferromanganese is used for reducing the iron oxide formed in the Bessemer's converter, and for counteracting the deleterious effects of phosphorus and sulphur. The resulting manganese oxides pass into the slag. Since ferromanganese contains about 4 or 6 per cent. of carbon, it is used for recarbonizing Bessemer's steel. Manganese steel is particularly hard and free from air blebs. It is used in the manufacture of burglar-proof safes, dredge pins, shoes and crusher plates for ore mills, etc. Cupro-manganese is an alloy of copper and manganese made by reducing the oxides of these metals in a graphite crucible or reverberatory furnace.

Heusler's alloy (1903) contains copper, aluminium, and manganese (55 : 15 : 30). They are magnetic. The aluminium can be replaced by arsenic, antimony, zinc, or tin, and the alloy remains magnetic.

Magnetic properties of the metals.—It is usual to regard iron, cobalt, and nickel as being particularly endowed with the magnetic quality, and Isaac Newton apparently believed that other bodies were quite uninfluenced by a magnet. Faraday and Tyndall, on the other hand, have said that it is doubtful if any substance is totally unaffected by a magnet. True enough, iron, cobalt, and nickel form a special group in that they acquire an enormous magnetization in comparison with other substances. Excluding the three elements just mentioned, there is no sign of any permanent magnetization with the other elements, and a feeble magnetic effect can be induced only with difficulty. The facts are here in agreement with the idea suggested by A. Schuster (1903): Every physical property hitherto discovered for one element has been found to be shared by all the others in varying degrees. While cobalt, nickel, iron, and Heusler's alloys are attracted by the poles of a magnet; graphite, bismuth, and several other bodies are repelled. M. Faraday (1845) called the former paramagnetic bodies (*παρά*, beside), and the latter diamagnetic bodies (*δίω*, across), because, when a rod of metal is suspended between the poles of a magnet, a paramagnetic body sets itself axially along a line joining the two poles, and a diamagnetic body sets itself equatorially, *i.e.* at right angles to the line joining the poles: Arranging the elements in two classes:

Paramagnetic elements: K, C, Ti, Ca, N, O, Cr, U, Mn, Fe, Co, Ni, Rh, Pd, Os, Ir, Pt
 Diamagnetic elements: H, Na, Cu, Ag, Au, Zn, Cd, Hg, Tl, Si, Sn, Pb, P, As, Sb,
 Bi, S, Se, Cl, Br

No connection between the chemical and magnetic qualities of the elements has been detected. Elements so much alike as potassium and sodium, oxygen and sulphur, nitrogen and phosphorus, titanium and silicon, are separated in different classes. The paramagnetic power of iron, cobalt, and nickel decreases with increasing atomic weight; while with phosphorus, antimony, and bismuth, and with copper, silver, and gold the paramagnetic power increases with increasing atomic weight. F. Heusler found that binary compounds of manganese—*e.g.* the nitride, MnN; phosphide, MnP; arsenide, MnAs; antimonide, MnSb; bismuthide, MnBi; silicide, MnSi; and stannide, MnSn, are also ferro-magnetic. These substances become non-magnetic at elevated temperatures; thus the antimonide becomes non-magnetic at about 310°, and the phosphide at about 18°. Hence, on a moderately cold day, manganese phosphide is magnetic, but it becomes non-magnetic when held in the hand.

Since all kinds of matter are affected by a magnet, and the effects produced differ only in degree, it follows that the atoms of the elements may all be magnetic so that magnetism is as inherent a property of the atoms as is gravitational weight. Still, the strongly magnetic characteristic of iron, and sometimes called ferromagnetism, is considered to be molecular and not atomic, because while in free iron, and in many iron alloys, the iron is ferromagnetic, yet in ferrous sulphate and many similar compounds it is paramagnetic, and in iron carbonyl and potassium ferrocyanide it is diamagnetic. According to W. Weber's theory, the molecule of every element is a magnet, but the element *en masse* usually shows no signs of polarity under ordinary conditions because the axes of the molecules point

in all directions at random, *A*, Fig. 188 ; but when a magnetic force is applied, some molecules are turned so as to face one way, *B*, Fig. 188. If the molecules persist in this position when the magnetic force is withdrawn, the body will be permanently magnetized ; and if the molecules return to

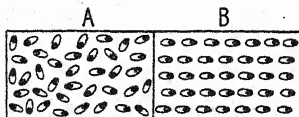


FIG. 188.—Diagrammatic Representation of the Molecules or Atoms.

their unoriented positions, magnetism will no longer be perceptible. There are many variations of this hypothesis.

Questions.

1. Given a solution of potassium permanganate, explain how the following substances can be obtained with its aid :—manganese dioxide, manganese sulphate, chlorine, oxygen.—*London Univ.*

2. Explain the term “available oxygen” as applied to potassium permanganate and potassium dichromate respectively ; also “available chlorine” as applied to bleaching powder.—*London Univ.*

3. What facts have led to the supposition that manganese is a heptad metal ? —*London Univ.*

CHAPTER XXVI

IRON, NICKEL, AND COBALT

§ 1. Iron—Occurrence, Preparation, and Properties.

History.—Iron is perhaps the most precious of all the metals, for civilized man would probably feel its absence more than would be the case if he were deprived of all the so-called precious metals. Early in the 14th century, when iron was scarce, some iron kitchen utensils in the household of Edward III. are said to have been classed among the jewels; and iron implements were among the most prized objects of plundering freebooters. The name iron is derived from the Scandinavian *iarn*. Several fabulous stories have been told describing how meteoric iron falling to the earth was sent from heaven as a gift of the gods to man. It requires no such romantic explanation of the discovery when it is remembered how readily iron is reduced from its ores. It has been suggested that "the first iron produced was the result of chance when lumps of iron ore, in place of stones, formed a rude cooking fire associated with some feast, where the fire was maintained long enough to effect the reduction." Then followed the observation that the higher temperatures obtained when the wind was blowing produced better material. Hence followed various contrivances for producing an artificial blast of wind, and so on by natural stages to the blast furnace. Iron implements have been used from prehistoric times, one was found during some blasting operations in the pyramid at Gizeh (Egypt), which is probably 5000 years old. The use of steel in China has been traced as far back as 2550 B.C., and we are told by the philologists that the early Vedic poets mention iron as being in the possession of their prehistoric ancestors, and that the artisans had acquired considerable skill in fashioning iron into tools. Owing to the fact that ancient objects of iron are comparatively rare, while ancient bronze objects are quite abundant, archaeologists affirm that the so-called *bronze age* preceded the *age of iron*. The argument is much weakened when it is remembered how much more readily iron is corroded by oxidation, and how much more readily bronze would survive atmospheric action. Still, bronze was more easy to extract and work than iron, and archaeologists consider that bronze ceased to be the dominant tool and weapon of civilized peoples about 500 B.C. The scarcity of copper and the abundance of iron in India makes it probable that with the Hindoos the iron age was not preceded by a bronze age. The Ayran emigrants who travelled into Europe carried with them a knowledge of producing iron from its ores. The Etruscans of North Italy were of Ayran stock, and they acquired considerable skill in the

working of iron. Soon after the downfall of the Roman Empire, the manufacture of iron developed in Spain. The famous Toledo steel blades were the product of Spanish artisans. Their fame spread, and Spanish craftsmen were drawn into France and Germany and there they introduced their peculiar Catalan forge which subsequently evolved into the large iron-smelting furnaces. The product of the Catalan forge is either a malleable iron or steel; the larger furnaces produced a variety of iron which could neither be forged nor tempered, although it was very suitable for all kinds of castings of moderate strength—cast iron. The discovery of a process by Cort—*vide infra*—whereby cast iron could be converted into wrought iron at a far less cost than was possible in the Catalan forge, gave a great impetus to the manufacture of iron in England.

In olden times, iron was symbolized by ♂, the spear and shield of Mars—the god of War—probably in allusion to its use in making weapons of war. Iron is frequently mentioned in the sacred writings. The Chalybes—an ancient nation living near the Black Sea—are supposed by the Greeks to have been the first to smelt iron ores. Hence the old term *chalybs* for steel, and our modern *chalybeate* for ferruginous. The process of smelting iron was early practised in the East, and the Hindoos acquired considerable skill in the manufacture of wrought iron and steel. The old sword blades of Toledo, Bilbao, and Damascus have never been excelled. The more recent method of smelting by means of the blast furnace is said to have been devised in Germany about 1350; and in Great Britain about 1500. Charcoal was first used as the reducing agent; in 1618 D. Dudley commenced using coal; and in 1713 Darby used coke. Coke and coal gradually displaced the use of charcoal. Some charcoal is still used where wood is cheap, *e.g.* in a few places on the Continent and in America.

Occurrence.—Small quantities of metallic iron occur in some basaltic rocks. An unusual mass, over 25 tons, has been found on the Disko Island, Greenland. Since iron rapidly corrodes when exposed to a humid atmosphere, native iron is not at all common. Nearly all *meteorites* or *aerolites* contain iron associated with other metals—chiefly copper, cobalt, and nickel. Traces of iron combined in various ways are found scattered almost universally throughout the mineral kingdom. Ferric oxide, Fe_2O_3 , is widely distributed in nature as *red hæmatite*—from the Greek *αἷμα* (*haima*), blood, in reference to the colour of its streak—*red ore*, and *specular iron ore*—from the Latin *speculum*, a mirror, in allusion to the lustrous crystals of this mineral. *Brown hæmatite* represents a class of hydrated oxides which may be represented by the general formula: $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$, where *n* represents the variable amount of water in different varieties—*limonite* is generally taken to be $\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$, that is, $\text{Fe}(\text{OH})_3$; *göthite*, $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$; and *bog iron ore* which occurs in Ireland belongs to this class. Limonite and scanthosiderite are considered to be colloidal forms of göthite with various quantities of adsorbed water. *Magnetite*, Fe_3O_4 , is called *loadstone*, and *magnetic oxide of iron*. *Siderite*, or *spathic iron ore*, FeCO_3 , is a ferrous carbonate. *Iron pyrites*, FeS_2 , and *chalcopyrites* or cupriferous pyrites, CuFeS_2 , are not worked directly for iron on account of the difficulty involved in eliminating sulphur from the product so that iron pyrites is often regarded as a sulphur ore. Very few clays, soils, and granite rocks are free from small quantities of iron. Iron plays an

important part in the nutrition of higher animals and plants, since this element seems necessary for their healthy growth.

Preparation and Properties.—Commercial iron always contains small quantities of graphite, iron carbide, iron phosphide, iron silicide, iron sulphide, and the corresponding manganese compounds. "Electrolytic iron" is made by the electrolysis of a salt of iron—ferric sulphate; and almost pure iron can be made by reducing a salt of iron—oxalate, chloride, or oxide—in a stream of hydrogen gas at 500° to 600° . If the reduction be carried out at a lower temperature, the black powder may become incan-

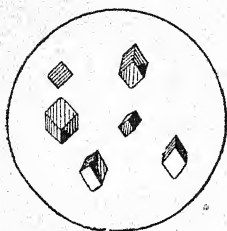


Fig. 189.—Cubic Crystals of Iron.

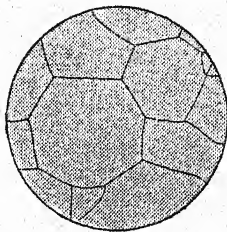


Fig. 190.—Ferrite.

descent on exposure to the air—pyrophoric iron. Pure iron crystallizes in the cubic system, Fig. 189. Iron is a grey, lustrous metal; it melts at 1500° and boils at 2950° . If a polished surface of a piece of iron be magnified about 150 diameters, a series of boundary

lines between the crystal walls appear as indicated in Fig. 190. The boundary lines are irregular because the crystals of the mass have been too closely packed together to enable them to develop their characteristic shape. For convenience, the pure metal is sometimes called ferrite.

Allotropic forms of iron.—If the temperature of a cooling bar of almost pure iron be recorded every half minute by a recording pyrometer the cooling process does not appear to be uniform and continuous, because the metal cools down to about 860° , and then becomes hotter; the cooling is then resumed until, at about 750° , the temperature again begins to oscillate. These temperature fluctuations are supposed to be due to the transition of iron from one allotropic modification to another. Each transition temperature corresponds with a change in the mechanical and physical properties of the iron. The sequence of changes is reversed when the cold iron is heated. Iron below 680° is called α -ferrite; between 750° and 860° , β -ferrite; and above 860° , γ -ferrite. Some consider that the alleged β -iron is a solid solution of γ -iron in α -iron. If the iron contains some carbon in solution, both the transition points approach 720° . The proximity of these points to 720° depends on the amount of carbon in solution. A new disturbance then appears in the cooling curve at about 660° . There is a marked evolution of heat at this temperature, for the red-hot cooling steel glows more brightly than before. This phenomenon was called by its discoverer—Barrett (1874)—the *recalcescence* of steel. Some differences in the physical properties of the three allotropic states of iron are summarized in Table XXXIII. The slope of the curve representing the absorption of hydrogen by iron is steepest between about 600° and 900° , and this is taken to be an effect of the changes in the character of the metal. 100 grams of iron at 416° absorb 0.035 mgrm. of hydrogen; at 878° , 0.315 mgrm. of hydrogen (or 1.58 mgrm. of nitrogen); at 1136° , 0.632 mgrm. of hydrogen (or 18.85 mgrm. of nitrogen). If

TABLE XXXIII.—COMPARISON OF THE PROPERTIES OF THE ALLOTROPES OF IRON.

Allotrope.	Normal limits of stability.	Magnetic properties.	Hardness.	Ductility.
α -iron	— to 750°	Magnetic	Soft	Ductile
β -iron	750° to 860°	Non-magnetic	Hard	Brittle
γ -iron	860° to —	Non-magnetic	Hard	Ductile

the iron contains carbon, some methane is formed at 850°–900°, by the absorbed hydrogen.

The action of acids.—Iron dissolves in dilute acids; sulphuric acid furnishes ferrous sulphate and hydrogen; and hydrochloric acid furnishes ferrous chloride and hydrogen. With cold dilute nitric acid (specific gravity below 1.034) hydrogen is not evolved, but the acid is reduced to ammonia, and this reacts with the excess of nitric acid to form ammonium nitrate; with an acid of specific gravity 1.034 to 1.115, *ferrous nitrate* is the main product; and with an acid of greater specific gravity than 1.115, say 1.3, *ferric nitrate* is the main product. With concentrated nitric acid (specific gravity 1.45), the iron does not dissolve. The iron in contact with the concentrated acid appears to have changed, for it behaves differently from a piece of the same sample of iron which has not been in contact with concentrated nitric acid. The sample which has not been in contact with the strong acid will precipitate copper from copper sulphate solutions, lead from lead nitrate, and silver from silver nitrate; the other sample will not. The inert iron is said to be in the passive condition. Passive iron does not dissolve when dipped in dilute nitric acid. Other oxidizing agents, chromic acid, hydrogen peroxide, will make iron passive. Chemists are not yet agreed as to the cause of passivity, the general idea is that a thin film of oxide is formed on the metal by contact with the oxidizing agent. The passivity can be removed by scratching the surface of the iron, by heating it in a reducing gas, by strongly rubbing the surface, and by bringing the passive iron in contact with zinc while immersed in the dilute nitric acid. Other metals also exhibit passivity, e.g. cobalt, nickel, chromium, and bismuth.

The rusting of iron.—E. L. Blassett said that one of the greatest metallurgical problems of the day is to produce a non-corrosive surface on iron and steel by electrochemical methods so as not to obliterate or obscure the minute outline of the articles. When commercial iron is exposed to a humid atmosphere for a short time, it soon becomes covered with a reddish-brown film which is called *rust*. Iron rust seems to be an indefinite mixture which on analysis furnishes numbers which vary according to the age of the rust, etc. Rust usually contains ferrous oxide, ferric oxide, carbon dioxide, and water. Analyses show that rust is probably a mixture of ferric oxide, hydrated ferrous and ferric oxides, and basic ferrous and ferric carbonates. If the rust has been long exposed to the air, the amount of ferric oxide is relatively large, and the amounts of ferrous oxide and carbon dioxide small. Rusting is a complex process, and workers are by no means agreed on the simple facts. Dry iron in dry air does not rust, moisture must be present before rusting can occur.

Some deny, others affirm, that the presence of an acid and water are necessary. It is exceedingly difficult to free water and the surface of glass from carbon dioxide; and silicic acid can be dissolved from the glass vessels used and from particles of slag in the iron. However, where careful attention has been taken to eliminate the disturbing factors, the evidence seems in favour of the conclusion that the presence of an acid is necessary for rusting; that an acid is always present when the iron dissolves; and it is highly probable that pure iron does not undergo appreciable oxidation when exposed to pure water and to pure oxygen. Films of moisture frequently condense on the surface of iron exposed to the air, and the moisture holds carbonic acid and oxygen in solution. The mechanism of the "atmospheric rusting" of iron may then proceed according to the following scheme: An acid ferrous carbonate, $\text{Fe}(\text{HCO}_3)_2$; or a basic carbonate, $\text{Fe}(\text{OH})(\text{HCO}_3)$, is first formed. The ferrous carbonate in contact with oxygen is oxidized to basic ferric carbonate, $\text{Fe}(\text{OH})_2(\text{HCO}_3)$; or to $\text{Fe}(\text{OH})(\text{HCO}_3)_2$; or both. The basic ferric carbonate is then hydrolyzed by the water, forming ferric hydroxide, $\text{Fe}(\text{OH})_3$; and the ferric hydroxide is subsequently more or less dehydrated, forming ferric oxide. Ferric oxide is more or less hygroscopic, so that once rusting has started, at any point, subsequent corrosion is quicker because the ferric oxide helps to keep the surface of the iron adjacent to the rust spot moist. Several other hypotheses have been suggested, and the subject is still *sub judice*. Enamels, paints, and bronze powders are largely used for protecting the metal from corrosion. Electro deposits have not proved efficacious. There are a number of processes which transform the surface into black oxide of iron or into a phosphate of iron. F. S. Barff heated the iron red-hot in a closed vessel, and treated it with superheated steam so as to transform the surface into the black oxide. There are several improvements in which a film of oxide can be produced on the metal without interfering with its temper. In the method known as *coslettizing*—named after its inventor, T. W. Coslett (1906)—ferric phosphate is deposited on the surface by pickling the cleaned article in a solution of phosphoric acid at 100° . Methods have also been proposed for depositing zinc (*q.v.*) and manganese dioxide on the surface of the metal to be protected.

Atomic weight.—The combining weight of iron determined by the analysis of the chloride, bromide, and iodide, and the synthesis of the oxide furnish numbers between 55.84 and 56.23 (oxygen = 16); and 55.85 is taken to be the best representative value. This agrees with the atomic weight deduced from the vapour density of the volatile compounds of iron; and by Dulong and Petit's method of approximation—specific heat of iron, 0.116.

§ 2. The Manufacture and Properties of Pig Iron.

The blast furnace for iron is the most efficient metallurgical instrument that exists.—W. M. JOHNSON (1914).

Iron ore carrying the equivalent of 20 to 70 per cent. of iron is abundant in many places, but it is sometimes mixed with so many impurities—sulphur, phosphorus, etc.—that it is not suited for making iron. The Swedish ore gives a product which is much prized. The oxides—magnetite

and hæmatite—and carbonates are the sources of commercial iron. These ores frequently contain a certain amount of clay, and the ores are then termed *clay ironstone*. The so-called *blackband ironstone* is a ferrous carbonate contaminated with clay, and black coally matters. The clay ironstones are usually calcined or roasted by stacking the ore with a small quantity of fuel in heaps, in stalls, or in shallow kilns. In the former case, combustion is started at one point and allowed to proceed throughout the whole mass. The temperature of the smouldering mass is sufficient to drive off most of the moisture, and carbon dioxide, and burn the organic matter and some of the sulphur and arsenic. The ferrous oxide is also oxidized to ferric oxide. This prevents the early formation of a fusible

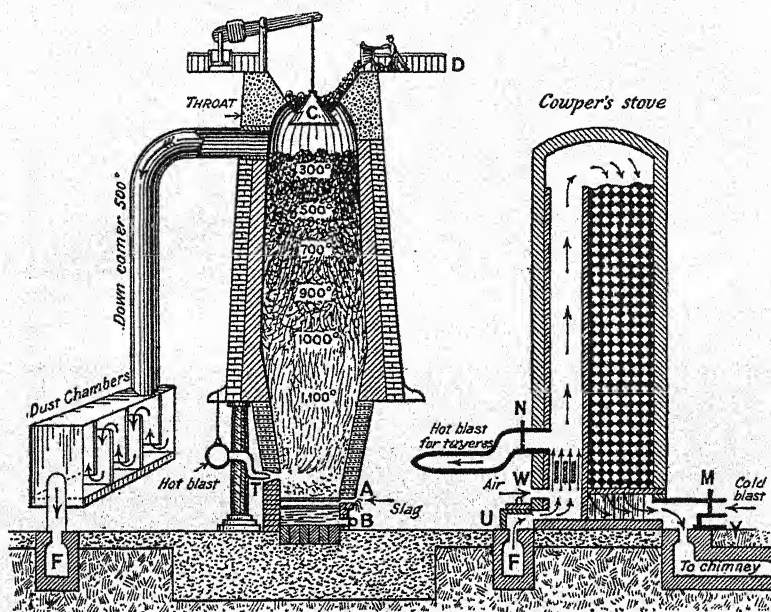


FIG. 191.—Blast Furnace (Diagrammatic).

slag which would attack the lining of the furnace. Ferric oxide does not form a slag at so low a temperature as ferrous oxide. At the same time the ore is made somewhat porous, and this facilitates its reduction to metallic iron at a later stage of the process.

The reduction of the calcined ore or of ferric oxide is effected in a blast furnace. The blast furnace is a long, vertical, cylindrical shaft fed with ore, fuel, and flux—all in lumps—at the top; and supplied with the air, under pressure, necessary for the combustion of the fuel at the bottom. The function of the blast furnace is to reduce the iron oxides to the metallic condition, and to free the iron so reduced from earthy admixtures by forming a slag which is fusible at the temperature of the furnace. The molten iron and slag are discharged at the bottom of the furnace, and the gases pass away at the top.

The blast furnace.—The smelting, as just indicated, is conducted in a tall cylindrical furnace—say, 80 feet high—and shaped approximately as indicated in the section, Fig. 191. The dimensions and constructional details of blast furnaces vary somewhat in different localities. The furnace now being described has an outer shell made of iron plates riveted together. Inside this is a casing of ordinary brickwork, and inside this a lining of firebricks— $\frac{1}{4}$ feet thick. The parts which are subjected to the greatest heat are built independently so as to facilitate repairs. The greatest internal width is 20 feet. This part is called the *bosh*. The mouth of the furnace is closed by a single, usually double, *cup-and-cone feeder*, *C*. The materials for charging the furnace are conveyed in trucks to the *charging gallery*, *D*, at the top of the furnace, and there tipped into the cup of the feeder. When the cup is filled, the cone is depressed, and the charge automatically distributed in the interior of the furnace. The waste gases pass away *via* the outlet at the throat of the furnace. The furnace narrows below the boshes, and at the hearth the diameter is 8 feet. Molten iron and slag collect on the hearth, and outlets are here provided, one, *A*, for tapping the slag, and another, *B*, for tapping the iron. Between 6 and 8 feet from the base of the furnace, six openings, *T*, are provided for the insertion of water-cooled nozzles—*tuyeres*—through which a blast of hot air is forced into the furnace.

The hot gases from the top of the furnace are led down a flue—the *down-comer*—into a chamber—the *dust-catcher*. The gases pass from the dust-catcher along an underground flue, *FF*, to a tower—*Cowper's stove*—packed checkerwise with firebricks. The flue gas is burnt in the combustion chamber of the stove, and the products of combustion pass on to the chimney. The secondary air required for the combustion of this gas enters through the ports, *W*. The burning gas raises the temperature of the checker brickwork. When the temperature of the stove is hot enough, the gases from the blast furnace are deflected, and burnt in an adjoining similar tower; meanwhile the gas and air valves—*U*, *V*, *W*—in the hot tower are closed; and another set of valves—*M*, *N*—connecting the tuyeres with the blowing machine are opened. The cold air passing through the hot checkerwork of the Cowper's stove on its way to the tuyeres is heated. When the tower has been cooled sufficiently, the adjoining stove is hot. The gas from the blast furnace is again burned in the cooled tower, and the blast is sent through the hot tower. Thus the towers are alternately heated by the combustion of the gas from the blast furnace, and cooled by the cold air from the blowing machine. In this way the blast of air is heated.

The reactions in the blast furnace during smelting.—The chemical changes which take place in the blast furnace during the smelting of iron ore are somewhat complex. Hence the following sketch must be regarded as a simplified description:

1. The ore.—The ore, mixed with coke and limestone, is exposed, in the upper part of the furnace, to the action of reducing gases, principally carbon monoxide, ascending from the lower part of the furnace. The action commences between 200° and 500° , that is, as soon as the charge has commenced its downward descent: $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightleftharpoons 2\text{Fe} + 3\text{CO}_2$; and reduction continues with increasing velocity as the charge descends into the hotter part of the furnace. There appears to be a complex series of side and intermediate reactions: $3\text{Fe}_2\text{O}_3 + \text{CO} = 2\text{Fe}_3\text{O}_4 + \text{CO}_2$; $\text{Fe}_3\text{O}_4 + \text{CO} \rightleftharpoons \text{CO}_2 + 3\text{FeO}$; and $\text{FeO} + \text{CO} \rightleftharpoons \text{CO}_2 + \text{Fe}$. Below dull redness, the back-reaction $\text{Fe} + \text{CO} \rightleftharpoons \text{FeO} + \text{C}$ is known to occur. Most of the oxide is reduced before it has descended 10 feet below the level of the charge; any oxide which has escaped reduction will then be reduced by the carbon: $\text{Fe}_2\text{O}_3 + 3\text{C} = 3\text{CO} + 2\text{Fe}$. The hot spongy iron meets the ascending carbon monoxide, and decomposes part: $2\text{CO} = \text{CO}_2 + \text{C}$. The solid carbon is deposited amidst the spongy iron. The iron undergoes little change until it reaches the zone of fusion. The iron, however, absorbs or dissolves much carbon as it passes down the furnace. The melting point of a mixture of iron and carbon is lower than that of pure

iron, so that while the temperature of the blast furnace would not be sufficient to melt pure iron to the necessary fluid condition, the temperature required for iron with carbon in solution is easily maintained. The molten iron trickles down and collects in the well of the furnace below the tuyeres. The iron takes up many other elements in addition to carbon during its descent in the furnace. Thus, silicon, sulphur, phosphorus, and manganese are partly absorbed by the iron, and partly by the slag.

2. The ascending gases.—The oxygen of the hot air blast burns the carbon of the hot coke: $C + O_2 = CO_2$; and the carbon dioxide is at once reduced by the hot carbon: $CO_2 + C = 2CO$. The ascending gases warm up the descending charge. When the temperature reaches about 600° , the limestone begins to decompose: $CaCO_3 = CO_2 + CaO$. Most of the carbon dioxide thus formed is at once reduced by the excess of carbon to carbon monoxide. At this stage, the reduction of the iron oxide to spongy metallic iron is practically complete. An excess of carbon monoxide is needed for the reduction because the reaction, $Fe_2O_3 + 3CO \rightleftharpoons 3CO_2 + 2Fe$, is reversible, and a condition of equilibrium would be attained when only a certain proportion of the ferric oxide is reduced. An excess of carbon monoxide favours a more complete reduction of the ferric oxide. There are quite a number of concurrent reactions taking place at the same time. If any water is present in the blast, it will be reduced: $H_2O + C = CO + H_2$; and the nitrogen of the air, brought in with the gas, forms a little cyanogen. The net result is a combustible gas, containing approximately:

CO	CO ₂	N	H	Hydrocarbons.
25·3	10·5	58·1	4·3	1·6 per cent.

The combustible gas is utilized for heating the blast; and if there be any surplus, it is used for heating the boilers which run the blowing engine; for calcining the ore, and for general heating purposes. If coal be used in place of coke, tar, etc., separate from the gas at the base of the down-comer and that mixture is treated by the process described under coal gas.

3. The slag.—When the charge in the furnace has descended about 20 or 30 feet, and the temperature is about 600° , it has formed a mixture of spongy iron, earthy gangue, coke, and limestone or quicklime. Little further change occurs until the temperature is hot enough to melt the mixture. At this temperature, a fusible slag is formed containing approximately 55 per cent. SiO_2 , 30 per cent. CaO , and 15 per cent. Al_2O_3 . The fused slag trickles down into the well, and floats on the surface of the molten iron. The slag is drawn from the furnace at intervals, and, when cool enough, tipped on the slag heap. The slag is derived from the ash of the fuel, from the earthy gangue of the ore, and the fluxes added with the charge to promote fusion. The composition of the slag varies with the quality of the iron, etc. Success in working the blast furnace depends largely upon the nature of the slag. Some types of slag which do not disintegrate on exposure to the air are used as road metal, and railway ballast; or mixed with some clay and moulded into bricks for paving, and building purposes. Some varieties of slag are made into cement.

The properties of pig iron.—The molten metal which collects in the well is tapped at intervals—say twice every twenty-four hours, and run into sand moulds, or into “chilled moulds,” and allowed to solidify. The iron is then called pig iron. The “pigs” of iron are about 3 feet long,

and 3 or 4 inches thick. In some cases the molten metal is run directly into a mixer, or into Bessemer's converter, and made into steel. Ore, flux, and coke enter the furnace; molten iron, molten slag, and gaseous products leave the furnace. Fresh charges of ore, coke, and flux are added in definite proportions at regular intervals, and the smelting of the ore thus continues without interruption for months or years.

The product from the blast furnace cannot be worked under the hammer, for when heated it passes immediately from the solid brittle condition to the liquid state in which it can be cast or poured into moulds—cast iron. The iron prepared in the older types of furnace—*e.g.* the Catalan forge—could not be melted in the industrial furnaces then in use, but when heated, it became soft and plastic so that it could be wrought under the hammer. Hence the principal varieties of iron are cast iron and wrought iron. The pigs of cast iron derived from the blast furnace are classed according to their quality. Ordinary pig iron contains from 1.5 to 4.5 per cent. of carbon. Higher proportions are sometimes present when the raw materials contain much manganese or chromium. The carbon of pig iron occurs in at least two different states—free and combined. Combined carbon is either in solution, or present as a definite chemical compound, see "Steel." The mode in which the carbon is associated with the iron has a marked effect on its properties. The free carbon, interspersed as graphite through the pig iron, is well shown on the fractured surface of a broken pig. If pig iron be digested with hydrochloric acid, the graphite remains behind as an insoluble black powder, but the combined carbon unites with the hydrogen forming various hydrocarbons which colour the solution yellowish-brown, and give the escaping gas a characteristic unpleasant smell. The gas is also contaminated with hydrogen sulphide, silicide, and phosphide derived from the impurities—iron sulphide, silicide, and phosphide—in the iron. The classification of pig iron is based on the relation of free to combined carbon. If much of the carbon be "free," the iron is called grey pig iron; and if much combined carbon be present, white pig iron; intermediate varieties are called mottled pig iron. The grey and mottled varieties are further subdivided. Pig irons too are often graded according to their source because certain districts work a specially pure or a specially foul ore, and this gives the iron from these districts characteristic properties. The following analyses will illustrate the difference between the three varieties of pig iron:

	Grey.	Mottled.	White.
Combined carbon (C.C.) . . .	0.90	1.80	3.00 per cent.
Free carbon (Gr.) . . .	2.8	1.40	0.10 per cent.

The pig iron also contains silicon, sulphur, phosphorus, and manganese. Cast or pig iron is not malleable, nor can it be welded. It is used for casting articles—like stoves and ornamental iron—which are not likely to be subjected to shocks. Cast iron is the starting-point for the manufacture of wrought iron and steel.

In the electrolytic purification of cast iron the metal is deposited in a rotating cathode, the electrolyte being a neutral solution of ferrous salts, which is circulated over iron turnings in order to maintain this neutrality. At regular intervals also additions are made to the solution of oxide of iron, which acts as a depolarizer and prevents hydrogen collecting in the cathode to a deleterious degree. The current density used is 1000 amperes

per square metre. The deposit obtained is of a high degree of purity, thus, starting from a cast iron containing 2.55 per cent. of C; 1.31 per cent. of Si; 0.07 per cent. of S; and 1.07 per cent. of P, the deposited metal contained 0.004 per cent. of C; 0.007 per cent. of Si; 0.006 per cent. of S; and 0.008 per cent. of P. In its unannealed condition, however, the iron is very hard and brittle, but on annealing for two hours at 900°, it shows the normal microstructure and is very tough under mechanical tests. The process is proposed for the manufacture of very thin tubing. The magnetic properties of the deposited metal, moreover, make it well suitable for the construction of transformers and other electrical appliances. For the making of high-quality cast steel it is claimed that the electrolytic metal should compete successfully with Swedish iron.

§ 3. The Manufacture and Properties of Wrought Iron.

Wrought iron is now made by melting pig iron with "scrap" iron on the bed of a reverberatory furnace, Fig. 192, lined with iron ore (Fe_2O_3). The reverberatory furnace—here called a *puddling furnace*—was first used for iron by T. and G. Granage in 1776, and more successfully by H. Cort in 1784, although similar furnaces had been used by copper smelters for some time. Rogers, in 1816, improved the furnace by introducing a bed of iron oxide in place of siliceous matters previously used. When the pigs of cast iron are melted in the puddling furnace, part of the carbon, silicon, sulphur, and phosphorus of the cast iron are oxidized by the furnace lining, and the metal melts to a fluid mass on the bed of the furnace (*melting stage*). The puddler then thoroughly mixes the charge so as to bring the molten metal into intimate contact with the iron oxide of the furnace bed. The puddler works the iron through an opening in the side of the furnace. The sulphur, phosphorus, and silicon are partly oxidized. Jets of flame soon appear on the surface of the molten metal—"puddler's candles" (*fluid stage*). The carbon is oxidized to carbon monoxide which burns to carbon dioxide. The other impurities are also oxidized, and form a slag with the lining of the furnace (*boiling stage*). The iron then becomes "pasty" because purified iron melts at a higher temperature than the less pure iron. The molten mass is stirred, puddled, and finally gathered into large "balls" or "blooms," each ball about 80 lbs. in weight (*balling stage*). The balls are removed from the furnace, and squeezed nearly free from slag by working under a steam hammer. The iron is then rolled into sheets so as to give the finished product a fibrous structure.

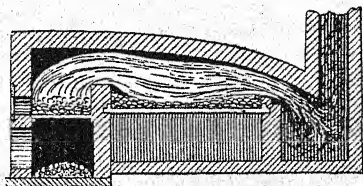


FIG. 192.—Puddling Furnace (Diagrammatic).

While cast iron melts at about 1200°, wrought iron melts at about 1550°. Wrought iron softens at about 1000°, and it can then be forged and welded. Wrought iron is tough and malleable, and fibrous in structure; cast iron is brittle, and it has a crystalline structure. Under the microscope, wrought iron appears to be composed of a bundle of fibres surrounded by some slag of magnetic oxide, etc.; each bundle seems to consist of a

series of fibres of metallic iron interlaced with a little ferrous silicide. These fibres give wrought iron its fibrous structure, and enable it to withstand severe longitudinal stresses. Wrought iron can be rolled into plates, and drawn into wire. It is made into wire, nails, chains, anchors, horseshoes, agricultural implements, etc. It is not used so much as formerly, because, on account of its high cost of production, it has been largely replaced by cheaper soft steels—indeed, some of the soft steels made by cheaper processes masquerade on the market as wrought iron. When heated red hot and quenched in cold water, wrought iron does not harden; steel under the same treatment becomes very hard.

§ 4. The Manufacture of Steel—Crucible and Cementation Processes.

The art of making steel is reduced to keeping iron in contact with carbon at a temperature sufficiently raised and maintained for a long enough time.—C. L. BERTHOLLET (1803).

The amount of carbon in steel is usually intermediate between that in cast iron and in wrought iron. Steel is made by decarbonizing cast iron, or by carbonizing wrought iron. The latter process gives the more reliable result because wrought iron is less contaminated with other impurities. The principal methods of making steel are: (1) the cementation process; (2) the crucible process; (3) electric processes; (4) Bessemer's acid and basic processes; and (5) Siemens and Martin's acid or basic processes.

The cementation process.—Bars of specially pure varieties of wrought iron—*e.g.* Swedish iron—are packed with charcoal in boxes made of fire-brick, and sealed with a lute of refuse from the troughs below the grind-stones of the steel grinders. The boxes are heated in a furnace for 8 to 11 days at about 1000° . The time and temperature depend upon the amount of carbon to be incorporated with the wrought iron. The bars, when removed from the cold furnace, have a blistered appearance, hence the term *blister steel*. The bars are broken and sorted by experts who estimate the quality from the appearance of fractured surfaces. The blistered steel is then heated, and hammered into bars. The product is a high-class tool steel called *shear steel*. The only change in composition which can occur during the cementation is due to carbon, and some of the purest steels in commerce are made by this process. Blister steel is also melted in crucibles and cast into ingots for high-grade cast steels. The cementation process is being gradually displaced by cheaper processes.

During cementation, solid carbon slowly diffuses into the iron. Probably at the high temperature, gaseous carbon compounds play some part in the action. Maybe carbon monoxide is formed and occluded by the iron. The carbon monoxide is then decomposed: $2\text{CO} = \text{CO}_2 + \text{C}$. The carbon dioxide escapes into the box of charcoal and there forms more carbon monoxide. This is again occluded by the iron, and the cyclic action: $\text{CO} \rightarrow \text{CO}_2 \rightarrow \text{CO} \rightarrow \dots$ is repeated indefinitely. Possibly also some cyanogen compounds take part in the action.

The crucible process for cast steel.—Bars of wrought iron are melted with a definite amount of carbon in fireclay crucibles. The iron slowly changes into steel by absorbing carbon. The time required is about four hours. Experience has taught the melter how much charcoal is needed to bring the metal up to the required carbon content. The success of the

operation depends upon the skill in the selection of the iron ; in the adjustment of the charge ; and on careful melting. Crucible steel is usually a high-grade tool steel used for razors, files, etc. A less pure product is made by heating a mixture of wrought iron with the necessary amount of cast iron. High-grade crucible steel is more expensive than Bessemer or open hearth steel. The term "cast steel" was originally reserved for crucible steel, but the cognomen is sometimes applied by vendors to steels made by cheaper processes. The manufacture of crucible steel was used by B. Huntsman in Sheffield about 1740 as a secret process.

Special steels of the self-hardening type are usually made by the crucible process, by alloying steel with small quantities of other metals ; these impart hardness, toughness, and strength. In districts where electric power is cheap, electric furnaces are coming into use for the manufacture of steel. Electric furnaces may not succeed in ousting the Bessemer and open hearth processes ; but they promise to play an important part in the future of the steel industry, and possibly may displace some of the older processes of making special steels.

§ 5. The Manufacture of Steel—Bessemer's Process.

In 1852 Kelly patented a process for purifying pig iron, based on the fact that if air be forced through a mass of molten pig iron, in a suitable vessel, the impurities—carbon, silicon, phosphorus, etc.—which prevent the pig iron being ductile and malleable are oxidized first, and a bath of molten metal, virtually wrought iron, is obtained. In 1856 H. Bessemer patented a converter suitable for the process. Bessemer afterwards bought Kelly's patents. The metal in the converter can be mixed with a known amount of spiegeleisen—i.e. a ferro-manganese containing a known amount of carbon. The steel is then at a sufficiently high temperature to permit of its being cast into moulds. This is really a wonderful process, for molten pig-iron at a red-heat is converted in a few minutes into molten steel at a white-heat, and the cost of the conversion is relatively small. The heat required to keep the charge molten and raise the temperature above the melting point of the metal is largely furnished by the combustion of the impurities. The steel can be prepared at but a small fraction of the cost of manufacture by the processes which precede. This method of making steel virtually revolutionized the art, but it will probably be ousted in a few years.

In Bessemer's process about 10 tons of molten pig iron are run into a large egg-shaped vessel, called the *converter*. The converter can be tilted into any required position. It is provided with holes at the bottom through which a powerful blast of air can be blown. The converter is made of wrought-iron plates, and lined with a bed made of silica and clay. A general idea of the structure of a converter can be gathered from Fig. 158, which shows part of the inside and part of the outside. Molten pig iron is run into the converter, and a powerful blast of air in fine jets is forced through the mass of molten metal. The temperature rises owing to the heat evolved by the oxidation and combustion of the impurities—the carbon, sulphur, and manganese. The carbon forms carbon monoxide which burns at the mouth of the converter. The flame is accompanied by a brilliant shower of sparks. The other oxides form a slag with the furnace

lining. Experience and the appearance of the flame tells the operator when to stop the blast. The right amount of spiegeleisen is then added to make a metal of definite composition. The blast is again turned on for a moment, and the metal is then cast into moulds to form blocks of Bessemer's steel.

Thomas and Gilchrist's basic process.—The sulphur and phosphorus are not removed by Bessemer's process just indicated. In 1878 S. G. Thomas and P. C. Gilchrist showed that, if the converter be lined with, say, dolomite (basic lining), and some lime be added to the charge of pig iron, and the blast continued a little longer, the oxides of phosphorus, sulphur, and silicon formed are absorbed by the furnace lining. The operation is otherwise conducted as before. The lining, after use, is called *Thomas' slag*, and it is used as a fertilizer on account of the phosphorus it contains. If the lining is siliceous, the operation is called the *acid Bessemer's process*; and if the lining be dolomite, the *basic Bessemer's process*.

§ 6. The Manufacture of Steel—Siemens and Martin's Open Hearth Process.

E. Martin (1864) made steel by fusing pig iron admixed with wrought iron scrap in an open sand basin; and W. Siemens (1863) by treating pig

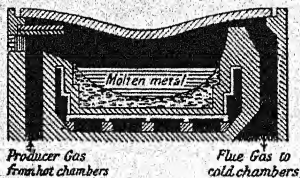


FIG. 193.—Hearth of Siemens-Martin's Process (Diagrammatic).

iron and pure hæmatite ore in a similar manner. The idea thus originated with the latter, though the process is usually called the "Siemens-Martin" process, or the open hearth process. In this process, the furnace is charged with a mixture of pig iron, scrap (wrought) iron, and good hæmatite ore free from carbon. The mixture is melted in a shallow rectangular trough or hearth. The furnace is heated by producer gas. Both the gas

and the secondary air for the combustion of the gas are pre-heated so that a very high temperature can be obtained. A general idea of the process can be gathered from Fig. 193, which shows a section through the hearth. The air port is not shown in the diagram. The gas and air burn on the left, the flue gases travel down the flue on the right, and in doing so heat up two chambers packed with checker brickwork. The direction of the burning gas is then reversed. Gas and air pass separately through the heated chambers, and the flue gases heat up another pair of similar chambers below the hearth. The direction travelled by the burning gas is reversed about every half-hour, and the heat of the flue gases is utilized in warming up chambers through which the unburnt gas and air will pass later on. The furnace is called *Siemens' regenerative furnace*. When a test shows that the metal contains the right amount of carbon, ferro-manganese is added as in the case of Bessemer's steel. If the bed of the furnace is made of siliceous materials—*acid process*—the proportion of carbon, silicon, and manganese are reduced during the treatment, but the amounts of sulphur and phosphorus remain fairly constant. In the *basic process*, the furnace is bedded with, say, dolomite, and there is a steady fall in the amount of

phosphorus and sulphur during the treatment, just as was the case with the basic Bessemer's process of Thomas and Gilchrist.

§ 7. The Constitution and Properties of Iron and Steel.

Steel is the mainspring of modern industry. The commercial importance of steel is greater than that of gold, silver, zinc, copper, and lead combined, and, indeed, the trite saying that this is an age of iron, is well founded.—W. M. JOHNSON (1914).

Pure iron is a laboratory curiosity. The properties of the commercial varieties of iron depend on the nature and proportion of impurities which are present. The foundry-man makes irons and steels of different composition to give them specific properties and to adapt them to certain uses. It may be said in a rough kind of way that steel is intermediate between cast and wrought iron so far as the proportion of carbon is concerned. Thus,

	Steel.				Wrought iron.
	Fig iron.	Hard.	Medium.	Soft.	
Carbon (per cent.)	3.0	0.9	0.4	0.1	0.06

but some of the so-called carbonless steels contain less carbon than wrought iron, and there are no hard-and-fast boundary lines. Indeed it does not seem possible to define steel satisfactorily without reference to its mode of manufacture. Some of the so-called carbonless steels have less carbon than wrought iron, but are distinguished from the latter by being molten when made. Iron at about 1240° dissolves 5 per cent. of carbon, the amount dissolved increases with the temperature, so that at about 2380° , nearly 10 per cent. is dissolved. A sample of iron of 99.98 per cent. purity melted at about 1520° ; and the freezing point of iron is lowered by the presence of carbon in the same way that the freezing point of water is lowered by salt (Fig. 59). The freezing point curve for solutions of carbon in iron is shown in Fig. 194. If a molten saturated solution of carbon in iron be slowly cooled, the excess of carbon separates as graphite, and the still molten mother liquid becomes poorer in carbon; as the temperature falls, graphite continues separating as illustrated by the curve *BO*, Fig. 194. When the mother liquid has 4.3 per cent. of carbon in solution, it solidifies *en masse* at 1130° , corresponding with the point *O*. The mixture containing 4.3 per cent. of carbon is the most fusible mixture of carbon and iron—the eutectic mixture, and 1130° is the eutectic temperature. If the cooling fluid contains less than 4.3 per cent. of carbon, a solid solution of about 2 per cent. of carbon commences to separate, and continues separating as the solution cools until the remaining fluid has the eutectic composition, which solidifies *en masse* at 1130° . The carbon which is rejected as the solution cools is called graphitic carbon. It is not attacked when the metal

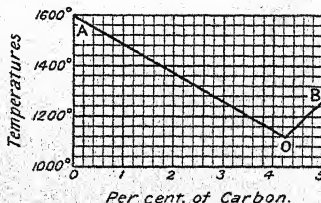


FIG. 194.—Freezing Curves of Carbon-Iron Solutions.

is digested in boiling hydrochloric acid, nor by nitric acid of specific gravity 1.2. When steel free from graphitic carbon is digested in dilute hydrochloric or sulphuric acid, (1) part of the carbon—called **hardening carbon**—appears to be combined with the metal, for it is evolved as a hydrocarbon gas along with the escaping hydrogen. J. Priestley called the gas “extremely foetid inflammable air.” (2) A portion of the carbon—called **cementite carbon**—remains undissolved. The cementite carbon, however, does dissolve in dilute nitric acid, forming a brown solution. The intensity of the colour is proportional to the amount of carbon dissolved, and by comparing the colour of the solution with the colour of the solutions made up with steels containing known amounts of carbon, the amount of carbon in solution can be estimated. This is the so-called *Eggertz's colour test for combined carbon in steel*.

Remember that iron above 860° is in the γ -condition; and between 860° and 750° in the β -condition. The solid solution of carbon in γ -ferrite

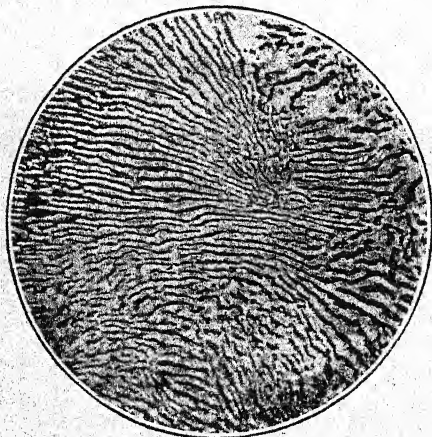


Fig. 195.—Pearlite.

above 860° is called **austenite**—after W. C. Roberts-Austen; and the solid solution of carbon in β -ferrite, **martensite**—after A. Martens. α -ferrite does not appear to form a solid solution of carbon in the same way as do β - and γ -ferrite. J. O. Arnold calls a solid solution of iron with 0.89 per cent. of carbon, **hardenite**. Hardenite has a constant composition 0.89 per cent. carbon. This may correspond with a definite carbide of iron, Fe_3C , or possibly with a solution of iron carbide, Fe_3C , in iron: $\text{Fe}_3\text{C} + 21\text{Fe}$. In any case, if hardenite be cooled

slowly, it decomposes just below 700° into a mixture of ferrite (iron) and cementite. There is a marked evolution of heat during the decomposition of the hardenite. This corresponds with the recalescence of cooling steel just below 700° . Cementite is an iron carbide, Fe_3C , or a solid solution containing Fe_3C with traces of several other carbides. The mixture of cementite and ferrite forms alternate layers of different degrees of hardness. When the surface of the metal is polished the harder parts stand out in relief. This gives the surface an iridescent appearance resembling mother-of-pearl when viewed at certain angles. Hence the mixture has been called “the pearly constituent of steel,” or simply pearlite. The alternate layers, under the microscope, have an appearance resembling that shown in Fig. 195. If the solid solution—hardenite—be cooled so quickly that the decomposition products of the hardenite have not time to segregate into alternate layers, but produce a more or less ill-defined mixture of cementite and ferrite, the mixture is

called sorbite—after H. C. Sorby. With slowly cooled steels the pearlitic structure is well developed; and with more quickly cooled steels, the sorbitic structure prevails. If the iron has less than 0·89 per cent. of carbon, ferrite separates from the cooling solid solution until a mixture of ferrite embedded in a matrix of hardenite with 0·89 per cent. of carbon remains; the hardenite then dissociates as indicated above. Similarly, if the solution has more than 0·89 per cent. of carbon, cementite separates until a matrix of cementite imbedded in hardenite, with 0·89 per cent. of carbon, is formed; the hardenite then dissociates as before. It will be noticed that pearlite and sorbite are not definite constituents; the terms refer to the mode in which certain constituents of a mixture are aggregated together. Metallographists are not all agreed in their interpretations of the observed phenomena, and details are still in the throes of disputation.

The sudden quenching of the steel while these changes are in progress may arrest or inhibit further change. Similarly, re-heating followed by sudden quenching may completely alter the structure, and accordingly also the properties of the metal. The hardness of steel thus depends upon its composition and its history—the way the metal has been cooled, and on its subsequent heat treatment. A typical steel containing nearly 1 per cent. of carbon when heated to a high temperature and suddenly chilled, becomes so hard that it will scratch glass, and so brittle that it will not bend very far without breaking. The operation is called hardening steel. The hardenite of hardened steel is preserved more or less undecomposed when the steel is suddenly chilled—hence the term “hardenite.” If the hardened steel be re-heated to its original high temperature, and slowly cooled—“letting down,” or annealing steel—it becomes soft and ductile. By re-heating hardened steel to certain definite temperatures—200° and upwards—and then cooling under definite conditions, steels can be obtained of varying, yet definite, degrees of hardness and ductility. This has been traced to changes in the constitution of the metal. The process of re-heating a hardened steel to a temperature far short of that employed when the steel was hardened is called tempering steel.

§ 8. Iron, Nickel, and Cobalt Monoxides.

Ferrous oxide, FeO .—Ferrous oxide is formed as a crystalline, black, magnetic substance when carbon dioxide is reduced by hot metallic iron. Ferrous oxide appears to be an intermediate stage in the reduction of ferric oxide by hydrogen or carbon monoxide, and hence the product is likely to be contaminated with either the higher unreduced oxide, or with metallic iron. Ferrous oxide is also formed when ferrous oxalate is heated out of contact with the air. When exposed to the air, ferrous oxide is oxidized to ferric oxide, Fe_2O_3 . The heat of formation is 60·4 Cals. White ferrous hydroxide, $\text{Fe}(\text{OH})_2$, is precipitated when an alkaline hydroxide or ammonia is added to a solution of a ferrous salt, provided air be entirely absent; but if the solutions have dissolved air, the precipitate will have a greenish colour. The white precipitate rapidly absorbs oxygen, and passes into ferric hydroxide, $\text{Fe}(\text{OH})_3$. Ferrous hydroxide and oxide dissolves in acids, forming ferrous salts. In acidic or neutral solutions, the ferrous salts do not oxidize quickly at ordinary temperatures, but in alkaline solutions ferrous iron is rapidly oxidized.

Cobaltous hydroxide, $\text{Co}(\text{OH})_2$.—A basic cobaltous chloride, $\text{Co}(\text{OH})\text{Cl}$, is formed as a blue precipitate when potassium hydroxide is added to a solution of a cobaltous chloride. The basic salt, on boiling, is converted into rose-red cobalt hydroxide, $\text{Co}(\text{OH})_2$, which, according to E. Donath, is slightly soluble in an excess of concentrated potash-lye forming a blue solution supposed to be *potassium cobaltite*, $\text{Co}(\text{OK})_2$. Copper oxide behaves similarly under the same conditions. Cobaltous hydroxide turns brown on exposure to the air owing to the absorption of oxygen. The hydroxide dissolves in hot concentrated potassium hydroxide, but it crystallizes from the solution on cooling. Cobaltous hydroxide thus shows feeble acidic properties. The hydroxide dissolves in ammonia, forming "cobalt-ammine" compounds (*q.v.*). The ammoniacal solution rapidly absorbs oxygen from the air. If cobalt hydroxide, carbonate, or nitrate be strongly heated in the absence of air to about 1000° , cobaltous oxide, CoO , is formed as a dark brown or olive-green powder. This oxide is stable in air, but when heated at about 600° – 700° in air it absorbs oxygen, and forms an oxide corresponding with cobalto-cobaltic oxide, Co_3O_4 . This in turn forms cobaltous oxide at about 970° . When heated in hydrogen or carbon monoxide, the cobalt oxides are reduced to metallic cobalt. Both the oxide and the hydroxide dissolve in acids, forming cobaltous salts.

Nickelous oxide, NiO .—When potassium hydroxide is added to a solution of a nickel salt, a pale green precipitate of nickelous hydroxide, $\text{Ni}(\text{OH})_2$ —approximately $4\text{Ni}(\text{OH})_2 \cdot \text{H}_2\text{O}$ —separates. Unlike ferrous and cobaltous hydroxides, this precipitate does not oxidize on exposure to the air. It dissolves in ammonia and ammonium salts, forming amines, and, unlike the corresponding cobaltous compound, the solution does not absorb oxygen from the air. If the hydroxide or carbonate of nickel be heated out of contact with air, a greenish powder of nickel oxide, NiO , is formed. The heat of formation is 51.5 Cals. It has been reported that a nickel suboxide, Ni_4O , is formed when nickelous oxide is fractionally reduced by hydrogen. This has not been confirmed. Nickelous oxide when heated in air forms nickel sesquioxide, Ni_2O_3 . Both the oxide and hydroxide dissolve in acids, forming nickelous salts.

§ 9. Iron, Cobalt, and Nickel Sesquioxides.

A voluminous reddish-brown precipitate of ferric hydroxide, $\text{Fe}(\text{OH})_3$, is formed when ammonia is added to a solution of a ferric salt. As in the case of aluminium hydroxide, $\text{Al}(\text{OH})_3$, there is some doubt whether a real hydroxide is formed. Several hydrated hydroxides are said to have been obtained by drying the precipitate at different temperatures, or by precipitating the hydroxide under special conditions; but these have not been proved to be true chemical individuals. Some of the hydrates occur in nature. For instance, *limonite* has a composition corresponding very nearly with $2\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$; *gothite*, $\text{FeO}(\text{OH})$ or $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$, isomorphous with diaspor and manganite. For "dialyzed iron," that is, a hydrosol of iron oxide, see p. 299. Ferric hydroxide is a very weak base, and the ferric salts are hydrolyzed in aqueous solution. The ferric oxide Fe_2O_3 occurs in nature as *specular iron ore*, which crystallizes in lustrous black crystals belonging to the hexagonal system; and in reddish masses of *haematite* isomorphous with corundum. Ferric oxide is a reddish-brown

powder formed when ferrous sulphate, or carbonate, or hydroxide, or many organic salts of iron are calcined in air. It is stable at a red heat, but at temperature above about 1300° it decomposes into Fe_3O_4 . The dissociation pressure at 1350° is nearly equal to that of oxygen in the atmosphere. Molten ferrous oxide in an atmosphere of oxygen freezes between 1562° and 1565° , forming ruby-red rhombohedral scales with a steel-like lustre, and a specific gravity 5.187 to 5.193. The powdered ferric oxide formed as just described is used as jeweller's polishing powder—rouge—and as a pigment. The particular tint of the pigment depends upon the temperature of calcination, which, in turn, appears to determine the size of the grains of the ferric oxide. Both ferric oxide and the hydroxide are basic, and give rise to ferric salts when heated with acids. Some of the native oxides, and the oxide which has been heated to a high temperature, dissolve in acids but slowly.

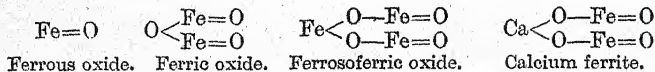
Nickelic oxide, Ni_2O_3 .—This oxide is said to be formed as a black powder when nickel nitrate or carbonate is ignited in air at a low temperature; but the oxides of nickel which have been described as Ni_2O_3 and Ni_3O_4 are usually regarded as mixtures of NiO_2 and NiO . Nickelic hydroxide, $\text{Ni}(\text{OH})_3$, is precipitated when chlorine is passed through water or alkaline hydroxide in which nickelic oxide is suspended; and also when a nickel salt is treated with a solution of bleaching powder. Nickelic oxide and hydroxide do not appear to be basic oxides, for when treated with acids, nickelous salts and oxygen or its equivalent are obtained: Thus $\text{Ni}_2\text{O}_3 + 6\text{HCl} = 2\text{NiCl}_2 + 3\text{H}_2\text{O} + \text{Cl}_2$; and $2\text{Ni}_2\text{O}_3 + 4\text{H}_2\text{SO}_4 = 4\text{NiSO}_4 + 4\text{H}_2\text{O} + \text{O}_2$. Hence nickelic oxide is a basic peroxide like manganese dioxide. When nickelic oxide reacts with ammonia, nitrogen is evolved and nickelous hydroxide is at the same time precipitated: $\text{Ni}_2\text{O}_3 + 2\text{NH}_3 + \text{O}_2 = 2\text{Ni}(\text{OH})_2 + \text{H}_2\text{O} + \text{N}_2$.

Cobaltic oxide, Co_2O_3 .—Cobaltic oxide is obtained by heating cobaltous nitrate or carbonate in air. It is a dark grey or brownish-black powder which forms an oxide corresponding with cobalto-cobaltic oxide, Co_3O_4 , when heated between the dissociation temperature 373° (760 mm.) and about 700° . Cobaltic hydroxide, $\text{Co}(\text{OH})_3$, is formed as a black precipitate when a cobaltous salt is treated with alkaline hypochlorite. Both the oxide and the hydroxide dissolve in acids, forming brown solutions which contain unstable cobaltic salts; these salts decompose when warmed, forming cobaltous salts and oxygen or its equivalent. Hence cobaltic oxide behaves as a feebly basic oxide and as a peroxide like nickelic oxide and manganese dioxide.

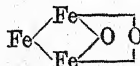
§ 10. The Higher Oxides of Iron, Cobalt, and Nickel.

Ferrosoferric oxide, or ferrous ferrite, Fe_3O_4 .—This substance—commonly called magnetite or magnetic oxide of iron—occurs free in nature in black octahedral crystals which are magnetic. It is the most stable oxide, and is formed when iron or iron oxides are heated in air or oxygen. According to Moissan, there are two modifications of ferrosoferric oxide: the one is formed by heating ferric oxide from 350° to 400° in a current of hydrogen or carbon monoxide; and the other by reactions at a high temperature, e.g. the combustion of iron in oxygen. The former is attacked by nitric acid, and has a specific gravity of 4.86; the latter is not attacked by the

same acid, and has a specific gravity 5.0 to 5.1. The phenomenon is thus analogous with the general effect of high temperatures on oxides like alumina, chromic oxide, ferric oxide, etc. Ferrosoferric oxide is not a basic oxide since it forms a mixture of ferric and ferrous salts when treated with acids. It is probably a "compound" oxide, analogous with manganomanganic oxide, Mn_2O_4 , and red lead, Pb_3O_4 . Molten ferrosoferric oxide freezes at 1527° . The relations of the iron oxides thus far considered will appear from the graphic formulæ:



According to this hypothesis, ferrosoferric oxide is a ferrous ferrite, $\text{Fe}(\text{FeO}_2)_2$. Among other formula proposed for magnetite there is



or a polymer of this. S. H. Enimens (1892) suggested that the presence of the atoms of iron attached to iron explain the magnetic qualities. Several other ferrites are known corresponding with the unknown ferrous acid HFeO_2 . Thus calcium ferrite, $\text{Ca}(\text{FeO}_2)_2$, that is, $\text{CaO} \cdot \text{Fe}_2\text{O}_3$, is made by precipitating a neutral solution of ferric chloride with lime water. Copper and cobalt ferrites, like ferrous ferrite, are strongly magnetic. Zinc ferrite, $\text{Zn}(\text{FeO}_2)_2$, is made by heating a mixture of the component oxides, and it is also formed in the extraction of zinc. In the wet process, zinc and ferric oxides are jointly precipitated in a solution containing ammonium chloride. Franklinite is a zinc ferrite with part of the iron and zinc replaced by manganese.

Cobalto-cobaltic oxide, Co_3O_4 , is formed in a similar way to ferrosoferric oxide, and it has probably the same constitution. The pure oxide has probably not been made. The oxide calcined at 750° has a little less oxygen than that corresponding with Co_3O_4 , and that calcined at about 1000° is virtually cobaltous oxide, CoO , a grey powder. There is some doubt about the existence of the corresponding nickelonickelic oxide, Ni_3O_4 , although it is reported to have been formed when moist oxygen is passed over nickel chloride heated to about 400° .

Dioxides.—Two probably isomeric nickel dioxides have been reported—a black dioxide is made by the action of bromine, or hypochlorites or hypobromites on nickelous hydroxide; and a green one by the action of hydrogen peroxide on nickelous chloride in the cold, followed by treatment with alcoholic potash. The green oxide is remarkably like hydrogen dioxide in its chemical properties, and is different from the black oxide. Hence it has been suggested that the green oxide is a peroxide or super oxide with nickel bivalent; and the black oxide is a dioxide or polyoxide, nickel quadrivalent:



If nickel oxide be heated with barium carbonate in an electric furnace, a dark coloured crystalline mass of barium nickelite, $\text{BaO} \cdot 2\text{NiO}_2$, is formed. Barium nickelite is decomposed by water. If cobalt hydroxide be

suspended in water, and then treated with hydrogen peroxide, the resulting liquid, after filtration, is strongly acid, and it is supposed to contain cobaltous acid, H_2CoO_3 . When a solution of cobalt sulphate is treated with iodine and sodium hydroxide, the resulting black precipitate has very nearly the composition CoO_2 —cobalt dioxide. A mixture of silver nitrate and potassium hydroxide can also be used as an oxidizing agent. The silver simultaneously precipitated is removed by washing the product with a solution of potassium permanganate and nitric acid. A green unstable solution of potassium cobaltite, K_2CoO_3 , is formed by the addition of potassium hydrogen carbonate, KHCO_3 , to the solution of cobaltous acid, and hydrogen peroxide, in the presence of sodium hydrocarbonate, gives an olive-green precipitate approximating to $\text{CoCoO}_2 \cdot \text{CoCO}_3 \cdot \text{H}_2\text{O}$, and thus consists of cobalt carbonate and cobaltate; copper sulphate similarly treated gives $\text{Cu}_2\text{CuO}_4 \cdot \text{CuCO}_3$.

Ferric acid and the ferrates.—When chlorine is passed through a concentrated solution of potassium hydroxide in which ferric hydroxide is suspended, the solution assumes a purple colour, and a black powder of potassium ferrate, K_2FeO_4 , separates. The reaction is usually represented: $2\text{Fe}(\text{OH})_3 + 10\text{KOH} + 3\text{Cl}_2 = 2\text{K}_2\text{FeO}_4 + 6\text{KCl} + 8\text{H}_2\text{O}$. The black powder appears to be analogous with potassium manganate, K_2MnO_4 , with potassium chromate, K_2CrO_4 , and with potassium sulphate, K_2SO_4 . The salt dissolves in water, forming a rose-red solution which, on crystallization, furnishes dark-red crystals isomorphous with potassium sulphate and chromate. Potassium ferrate is unstable and its solution readily decomposes: $4\text{K}_2\text{FeO}_4 + 10\text{H}_2\text{O} = 8\text{KOH} + 4\text{Fe}(\text{OH})_3 + 3\text{O}_2$. The barium salt, BaFeO_4 , is a dark carmine-red and fairly stable. By analogy with the chromates and sulphates, it is inferred that the ferrates are derived from an unknown ferric acid, H_2FeO_4 ; which in turn is derived from an unknown ferric anhydride, FeO_3 , analogous with sulphur and chromic trioxides.

§ 11. Cobalt and Nickel.

History.—The word “kobalt” occurs in the writings of B. Valentine and Paracelsus to denote a goblin supposed by the old Teutons and Scandinavians to haunt the mines, destroyed the work of the miners, and gave them a lot of unnecessary trouble. The word is derived from the German *kobald*, an evil spirit. The church service in some mining districts once included a prayer that God would preserve the miners and their work from kobalds and other evil spirits. The term was applied to what were called ‘false ores,’ that is, ores which did not give metals when treated by the processes then in vogue for the extraction of the metals; and also to ores which had an objectionable smell. The term was gradually confined to the minerals used for colouring glass blue, and which are still used for making smalt. In 1735 Brandt stated that the blue colouring principle is due to the presence of a peculiar metal or semi-metal which he called “cobalt rex,” hence our “cobalt.” Brandt isolated the metal in 1742.

Nickel seems to have been known to the Chinese in early times. In Europe, towards the end of the 17th century, the German term *kupfer-nickel* (false-copper) was applied to an ore which, while possessing the general appearance of a copper ore, yet gave no copper when treated by the general process then used for the extraction of copper. Never-

theless the mineral was supposed to be an ore of copper, or a species of cobalt or arsenic united with copper. A. F. Cronstedt, 1751-1754, stated that kupfer-nickel contains a metal which gives a brown, not a blue colour, with glass. Cronstedt's views were not accepted generally until T. Bergman had shown clearly that Cronstedt's metal was a new element in an impure condition. Bergman's arguments were mainly directed to controvert the view that nickel is a compound of cobalt, iron, arsenic, and copper, advocated by le Sage. He showed that nickel retains its individuality when arsenic is absent, and when no signs of cobalt or copper can be detected in solutions of the metal. Bergman also tried unsuccessfully to make nickel synthetically from mixtures of the elements in question. He was unable to prepare the metal free from iron, and consequently was led to say "there are indeed many and weighty reasons which induce us to think that nickel, cobalt, and manganese are perhaps to be considered in no other light than modifications of iron." He then applied an argument not unlike that employed by Davy in the case of chlorine, and which is worth repeating :

We must be carefully on our guard against such glimmering lights which are generally fallacious. So long as no one is able to produce these bodies from pure iron, and to explain in a clear and intelligible way the process by which manganese, cobalt, or nickel may at pleasure be generated, such vague suspicions must give way to phenomena which are constants and the substances themselves continue to be considered of an origin altogether distinct and peculiar to themselves.

Occurrence.—Cobalt and nickel are nearly always found associated with one another. Both elements occur free in some meteorites. Cobalt occurs as a minor constituent in some minerals. It also occurs combined as arsenide in *smaltite* or cobalt speiss, CoAs_2 ; and as *cobaltite* or cobalt glance, CoAsS . Nickel occurs as *kupfernickel* or *niccolite*, NiAs ; *millerite*, or nickel blende, NiS ; *nickel glance*, NiAsS ; and *garnierite*, a silicate of magnesium and nickel, $(\text{NiMg})\text{H}_2\text{SiO}_4$, found in New Caledonia.

Preparation.—The cobalt ores are usually worked to get cobalt salts, and not the metal. The ores are first roasted to remove arsenic and sulphur. The resulting oxides are digested with hydrochloric acid, and the solution treated with hydrogen sulphide to remove the copper, lead, antimony, etc. Bleaching powder is then added in just sufficient quantity to oxidize the iron, which is then precipitated by the addition of chalk. The clear solution is treated with more bleaching powder to precipitate the cobalt oxide; any nickel which might be present is precipitated by the addition of milk of lime. Cobalt metal is obtained by reducing the oxide in a current of hydrogen, or by reducing the oxide with aluminium powder as in Goldschmidt's process for chromium, Fig. 187.

Nickel and cobalt ores are treated by different methods depending upon the nature of the ore under treatment. Many ores are treated by a process similar to that described for the extraction of copper. The ore is "Bessemerized" in a converter, Fig. 158, so as to oxidize the iron, and furnish a matte rich in copper and nickel. This is roasted, and the resulting oxides are reduced with coke. An alloy of copper and nickel is thus obtained. The two metals are separated by an electrolytic process.

A nearly pure nickel is made by Mond's process, which is based upon the formation of a readily volatile compound of nickel—nickel carbonyl, $\text{Ni}(\text{CO})_4$. The roasted ore (oxides) is heated in a reducing atmosphere to

about 300° so as to transform the oxides to metal. The resulting mixture is heated with carbon monoxide under a pressure of about 15 atmospheres at a temperature of about 100° . The vapours of the escaping nickel carbonyl can be decomposed by heating the gas to 200° under atmospheric pressure. The escaping carbon monoxide is used again.

Properties.—Cobalt and nickel are hard white metals, cobalt is slightly bluer than nickel. Both metals are malleable and ductile. A sample of nickel of 99·84 per cent. purity melted at 1452° , and a sample of cobalt of 99·95 per cent. purity at 1478° . Iron is strongly magnetic, cobalt and nickel are but feebly magnetic; iron forms Fe_3O_4 when heated with steam; nickel and cobalt form the monoxides. Iron rusts in moist air, cobalt and nickel only oxidize very slightly unless heated in air. Dilute hydrochloric and sulphuric acids dissolve cobalt and nickel slowly, whereas iron is rapidly dissolved by these acids. Cobalt and nickel are both attacked by nitric acid, forming the corresponding nitrates— $\text{Co}(\text{NO}_3)_2$ and $\text{Ni}(\text{NO}_3)_2$.

Alloys and Uses.—Nickel is much used for “nickel plating” other metals on account of its silvery appearance, and the fact that it does not readily tarnish in air. The nickel is deposited from a double sulphate of ammonium and nickel by a process similar to that used for “silver and gold plating.” Iron can be deposited on copper—the so-called *steel-plating*—from a solution of ferrous and ammonium sulphates in the proportion of about 50 : 1. Nickel is used in making several important alloys: *E.g.* *German silver* is the trade-name for a series of white alloys consisting of a mixture of copper, nickel, and other metals. The standard might be regarded as: copper 50; nickel 25; zinc 25 per cent. Additions of lead, tin, cadmium, and iron are used in different proportions to suit various uses to which the alloy may be put and to accentuate maybe hardness, ductility, surface polish, electrical conductivity, etc. Naturally several other trade-names are used—*nickel silver*, *silveroid*, *argentoid*, *nickeline*, etc. *Manganin* is an alloy used in making electrical resistances. It is a kind of German silver with manganese in place of zinc—*e.g.* manganese 12, copper 85, nickel 3. *Nickel steel* is hard and tough, and is used for parts of machinery designed to withstand continuous wear and shocks, and in the manufacture of armour plates, burglar-proof safes, etc. A nickel steel containing about 35 per cent. of nickel has a very small coefficient of thermal expansion, and is consequently useful for making measuring instruments, pendulums, etc. It is called *invar*; an alloy with about 40 per cent. of nickel expands at about the same rate as glass, and hence it can be sealed into glass in the manufacture of incandescent electric lamps, etc. This alloy is called *platinite*. Nickel coins contain about 25 per cent. of nickel, and 75 per cent. of copper. An alloy of nickel 60, copper 33, is called *monel metal*. It is said to be less liable to corrode than brass or bronze and nearly as strong as steel. An alloy containing 80 to 95 per cent. of copper and 5 to 20 per cent. of a cobalt-tin (2 : 3) alloy is one of the so-called *non-corrosive alloys* because it resists fairly well the attack by many acids which rapidly corrode copper. Finely divided nickel is used as a catalytic agent since, in the presence of hydrogen, it is able to reduce or hydrogenize many organic and inorganic compounds. Traces of sulphur or chlorine in the gas rapidly makes the nickel lose its activity.

✓ Cobalt oxides and silicates are used for colouring glass, and pottery

glazes, etc. Smalt is a glass made by fusing cobalt oxide and silica, and when ground it is used as a pigment. *Thénard's blue* is a blue pigment made by calcining a mixture of cobalt oxide with alumina or aluminium phosphate or arsenate. *Thénard's blue* is usually regarded as a *cobalt aluminate*, $\text{Co}(\text{AlO}_2)_2$; and *Rinnmann's green*, as a *cobalt zincate*. The two colours, however, are developed with binary mixtures in very different proportions, and accordingly the colour can be regarded as a solid solution of, say, cobalt and zinc oxides, or of $\text{CoO} \cdot \text{ZnO}$ in an excess of one of the other oxides. Magnesia and cobalt oxide treated in the same way furnish a red colour.

Atomic weights of cobalt and nickel.—The almost identical values for the atomic weights of cobalt and nickel has attracted much attention. Nickel oxalate, cyanide, sulphate, chloride, bromide, iodide, etc., have been investigated, and the result shows that the combining weight of nickel (oxygen = 16) lies somewhere between 58.03 and 58.95; and 58.68 is generally accepted as the best representative value. Similarly for cobalt the combining weight lies somewhere between 58.7 and 59.7; and 58.97 is generally taken to be the best representative value. These numbers agree with the atomic weights estimated from the specific heats by Dulong and Petit's approximation.

§ 12. The Relationships of the Members of the Iron Family

The object of a classification of chemical substances is the arrangement of them in such a way that the position in the system of each substance may express its own chemical nature and the relation in which it stands to other substances.

The atomic weights of the iron, cobalt, and nickel group are not very different numerically. It is not at all uncommon to find that elements with but small differences in their atomic weights show marked differences in their physical and chemical properties. For example, boron, 11, and carbon, 12; potassium, 39.1; argon, 40; calcium, 40.1; selenium, 79.2; and bromine, 79.96; iodine, 126.97; tellurium, 127.6. The physical properties of the iron, cobalt, and nickel family are summarized in the table:

TABLE XXXIV.—PHYSICAL PROPERTIES OF THE IRON FAMILY.

	Iron.	Cobalt.	Nickel.
Atomic weight	55.85	58.97	58.7
Specific gravity	7.8	8.5	8.8
Atomic volume	7.16	6.94	6.68
Melting point	1520°	1490°	1452°
Boiling point	2450°	—	2450°

The elements are usually associated together in nature; they are all magnetic, nearly white, hard metals with a high melting point. Their chemical properties exhibit a gradual transition from iron to nickel. Iron forms two well-defined basic oxides, so does cobalt, but the basicity of the cobalt sesquioxide is so feeble that the corresponding salts are only known in solution—the double salts, however, are stable. Nickel gives only one

basic oxide—nickel monoxide—the sesquioxide does not appear to form salts. Although the atomic weight of nickel is less than that of cobalt, the physical and chemical properties of the members of the family show a transition from iron to cobalt to nickel. This is not a common phenomenon with the other family groups. It is common to find that in groups of related elements each member seems to affect a particular state of oxidation or combination in which it is in a condition of maximum stability, or chemical repose, as W. A. Tilden expresses it (1895), and to further emphasize the idea, he points out that the aluminic salts cannot be reduced to a lower state of oxidation; chromic salts can be reduced with difficulty; ferric salts are easily transformed to ferrous salts; while the manganic salts can be reduced by mere heating. Ferric chloride is a stable salt; cobaltic chloride is stable only in solution; and nickelic chloride is unknown. A series of polyiodides, MI_2I_2 or MI_4 , are known—where M stands for an atom of manganese, iron, cobalt, or nickel. The special property of cobalt in forming well-defined ammines is worthy of note.

When an element occurs in two or more states of oxidation as in ferric and ferrous iron, the compounds of each behave as if they were derived from two different elements, and the derivatives of each are related to two totally different families of elements. Thus, the salts of trivalent or ferric iron are related to the salts of trivalent aluminium and chromium; while the salts of bivalent or ferrous iron are related to the magnesium or zinc compounds. In neither case does the element corresponding with these analogies—say aluminium in the one case and zinc in the other—bear any marked resemblance to metallic iron. Any theory of atomic structure must bring out the dual, or complex personality of the elements. The three elements are related to aluminium, manganese, and chromium through iron (ferrates, chromates, manganates, and the alums); to the zinc and magnesium family through ferrous iron; and to copper through nickel. The isomorphic relationship of the elements iron, cobalt, nickel, aluminium, manganese, and chromium is not only emphasized by the crystalline form of many salts but also by their constant association in native minerals—*e.g.* the replacement of aluminium by iron in silicates. Nontronite, $Fe_2O_3 \cdot 2SiO_2 \cdot 2H_2O$, and kaolinite, $Al_2O_3 \cdot 2SiO_2 \cdot 2H_2O$, for instance, appear to be the terminal members of a series with all intermediate proportions of ferric and aluminium oxide, and there are reasons for supposing chromic oxide can act in a similar way. The contrast of chromic iron ore, $FeO \cdot Cr_2O_3$, with magnetic oxide of iron, $FeO \cdot Fe_2O_3$, is also of interest. To emphasize the relation between copper and nickel, the general colour of their salts; the solubility of the hydroxide in an excess of ammonia to form blue solutions; the isomorphism of their two sulphates in spite of the fact that their *stable* forms are $CuSO_4 \cdot 5H_2O$ and $NiSO_4 \cdot 7H_2O$ respectively. The two sulphates form mixed crystals such that if the copper be in excess, both copper and nickel form pentahydrates, and if the nickel be in excess, heptahydrates.

Questions.

1. State the essential chemical difference between the Bessemer and the open-hearth processes for the production of steel. What changes are common to both processes? What is the "basic Bessemer process," and why is it of importance?—*Worcester Polytechnic Inst., U.S.A.*

2. Give an account of the metal nickel, and contrast its properties with those of allied elements.—*London Univ.*

3. Why should the limestone in the Bessemer and the Thomas Gilchrist processes purify steel, when limestone is always present in the previous blast furnace process?—*Amherst Coll., U.S.A.*

4. Write the equation for the action of potassium permanganate on ferrous sulphate and dilute sulphuric acid.

5.—(a) How would you make sodium chromate from chromium sulphate? (b) Show how you would change the product above into chromium sulphate and write the equation.—*Amherst Coll., U.S.A.*

6. How would you make (1) ferrous chloride from iron, (2) ferrous sulphide from ferrous chloride, (3) ferrous sulphate from ferrous sulphide, (4) ferric hydroxide from ferrous sulphate? Write the equations for the reactions after giving the descriptions of the processes and indicate the colour of the product in each case.—*Sheffield Scientific School, U.S.A.*

7. Describe briefly and explain the blast furnace process, giving the purpose of each ingredient of the charge.—*Amherst Coll., U.S.A.*

8. What is the difference in composition between wrought iron, cast iron, and steel?—*Sheffield Scientific School, U.S.A.*

9. How is ferrous chloride converted into ferric chloride? How is ferric chloride changed to ferrous chloride?—*Sheffield Scientific School, U.S.A.*

10. What is the difference between a blast furnace and a reverberatory furnace, and what is the chief use of each? What is meant by a "flux," and what is one of the most commonly employed substances of this class?—*Princeton Univ., U.S.A.*

11. What is the difference in chemical composition between cast iron, wrought iron, and steel? How and under what conditions do the following substances act upon iron: Water, copper sulphate, sulphuric acid, aqua regia, chlorine?—*London Univ.*

12. Describe exactly how you would prepare ferrous sulphate and ferric chloride from metallic iron, and also how you would prepare ferrous phosphate from the first, and ferric oxide from the second of these salts; state the appearance of each product.—*London Univ.*

13. What are the chief sources of nickel? How can nickel be obtained free from cobalt? For what purpose does nickel find employment in the arts? By what properties are nickel compounds distinguished from those of cobalt?—*Board of Educ.*

14. What are, most probably, the highest states of oxidation of chlorine, chromium, iodine, and iron? Give an account of the experimental evidence in favour of the views you support.—*Board of Educ.*

15. What proportion of carbon is necessary in mild steel and hard steel respectively? and what views are now held as to the condition in which carbon exists as steel?—*Board of Educ.*

16. Would the following equation represent correctly the formation of potassium ferrate from ferrous hydroxide when the latter is suspended in a solution of potassium hydroxide, and chlorine afterwards passed into the solution: $\text{Fe}(\text{OH})_2 + 2\text{KOH} + 2\text{Cl}_2 = \text{K}_2\text{FeO}_4 + 4\text{HCl}$? What objections can be urged against the equation as describing the formation of potassium ferrate in this way?—*R. Galloway.*

17. What is meant by the statement that the combining weight of hydrogen is 1, that of oxygen 8, and that of iron 28? How are these values arrived at? Why is the combining weight doubled to give the atomic weight of either oxygen or iron?—*Board of Educ.*

18. Describe, with all essential practical details, the preparation of pure specimens of the following salts:—(a) ferrous sulphate from iron pyrites, (b) mercuric chloride from cinnabar, (c) barium nitrate from barium chloride.—*Board of Educ.*

19. What are white, blue, and green vitriols? Why does ferrous sulphate become brown on exposure to air?

20. A 6 per cent. aqueous solution of ammonium nickel sulphate, $(\text{NH}_4)_2\text{SO}_4 \cdot \text{NiSO}_4 \cdot 6\text{H}_2\text{O}$, is employed for the electro-deposition of nickel. How much pure nickel, 30 per cent. sulphuric acid, pure ammonium sulphate, and water are theoretically needed for making 12 kgms. of solution.—*G. Bornemann, Stöchiometrie, Leipzig, 1909.*

CHAPTER XXVII

THE OXYGEN COMPOUNDS OF NITROGEN

§ 1. Sodium and Potassium Nitrates.

POTASSIUM and sodium nitrate are two important salts. The former is also called "nitre" and "saltpetre." The word "nitre" is derived from the Arabic *nitrum* or *natrum*, thence the Greek words *nitron* or *natron*, meaning "soda," while saltpetre is a corruption of the Latin *sal petrae*—"salt of the rock." The salt designated by these two terms is very different from rock salt, and from soda. Hence, in one of Solomon's sharpest proverbs (Prov. xxv. 20), the reference to the commotion which ensues on mixing vinegar with nitre shows that what we call sodium carbonate was in the writer's mind. Sodium nitrate is often called "soda nitre" or "Chili saltpetre," to distinguish it from "potash nitre" or "saltpetre" proper. Nitre was probably known in India and China before the Christian era.

Crystallization of the two nitrates.—Potassium nitrate forms hexagonal prisms (rhombic system), and sodium nitrate rhombohedral

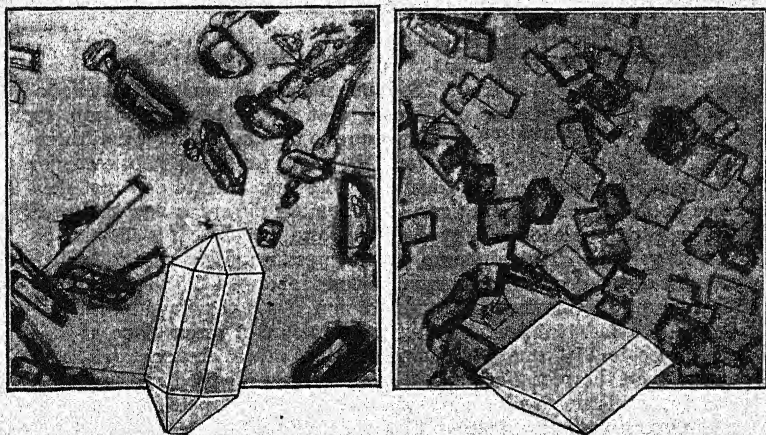
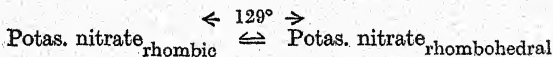


FIG. 196.—Potassium Nitrate (left) and Sodium Nitrate (right) Crystals.

(trigonal system) crystals. The crystals are illustrated in Fig. 196, where the salts have been crystallized on a slip of glass and photographed

under the microscope. Ideal crystals are illustrated by the outline drawings. The form of the crystals of sodium nitrate explains why it is sometimes called *cubic nitre*. It is really interesting to watch the crystallization of a drop of warm, slightly supersaturated solution of potassium nitrate on a glass slip under the microscope. Crystallization starts at the edges. Here rhombohedral crystalline plates (left, Fig. 196), which are not really isomorphous with the crystals of sodium nitrate (right, Fig. 196), are first formed, although a mixture of the two salts is formed; these are quickly followed by needle-like rhombic (trigonal) crystals resembling calcite. As a matter of fact, both forms of crystals appear in the photograph (left, Fig. 196). Immediately the rhombohedral crystals touch the rhombic crystals, the latter lose their sharp outlines, and needle-like rhombohedra sprout forth on all sides. Hence, *potassium nitrate is dimorphous*. The rhombic crystals are unstable above, and stable below 129° ; and conversely, the rhombohedral crystals are stable above, and unstable below 129° . Hence, 129° is a transition temperature:



Both potassium and sodium nitrates are soluble in water; and, although sodium nitrate is more soluble than potassium nitrate at ordinary temperatures, the reverse is true at 100° . This is illustrated by the solubility curves—Fig. 197—(grams of salt per 100 grams of solvent) from 0° to 100° . The solubilities of the nitrates of the alkali metals and ammonium nitrate, expressed in grams per 100 grams of solution at 20° :

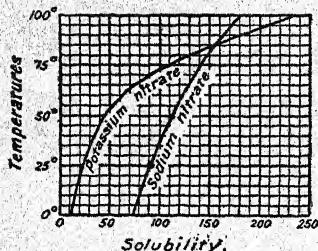


FIG. 197.—Solubility Curves of Sodium and Potassium Nitrates.

Cs	Rb	K	NH ₄	Na	Li
18.7	34.6	24.0	65.8	46.7	41.2

Sodium nitrate is somewhat deliquescent, that is, it becomes damp on exposure to the air, and it cannot, therefore, be used for some purposes for

which potassium nitrate is applicable. Potassium nitrate does not deliquesce under the same conditions. Ammonium nitrate is now preferred to the potassium salt for some types of gunpowder because it is cleaner and leaves no residue. The minimum vapour pressure of water at 20° which causes the deliquescence of ammonium nitrate is between 8.5 and 10.8 mm.; of sodium nitrate, between 12.3 and 13.5 mm.; and of potassium nitrate, between 15.5 and 16.5 mm. The vapour pressure of water alone at this temperature is 17.4 mm. Sodium nitrate is much cheaper and more plentiful than the potassium salt. Sodium nitrate is used in the manufacture of fertilizers; and of sulphuric acid, nitric acid, potassium nitrate, sodium arsenate, fireworks, glass, etc.; in curing meat; etc.

The occurrence of sodium nitrate.—Large quantities of sodium nitrate occur in the Argentine, California, and principally in the rainless zone on the West Coast of South America—Peru, Bolivia, Chili. The salt occurs in large flat basins between the ridges on the Tarapacca plateau.

Fig. 198 represents a diagrammatic cross section through a "nitre basin" which will give a rough idea how soda nitre occurs: There is first a 2 or 3 inch surface layer of grey sand and pebbles. The surface itself is almost devoid of vegetation. Below the surface is a 1 to 5 ft. layer of similar material cemented together with clay and salt and sodium nitrate. This stuff is called by the natives "costra." Below the "costra" is a white



FIG. 198.—Geological Section of Nitre Bed (Diagrammatic).

rock of shale, or limestone, or other rock which may be there outcropped. Costra is a kind of low-grade nitre rock or caliche running 5 to 12 per cent. sodium nitrate which does not pay to work; the caliche runs 18 to 25 per cent., and in exceptional cases 50 or 60 per cent. sodium nitrate—the average runs 20 to 30 per cent. The deposits are close to the surface, and naturally vary a little in different places.

The extraction of sodium nitrate.—The caliche is mined by boring down to the lowest stratum. The bore is enlarged and charged with blasting powder. The explosion breaks up the nitre bed within a 50 ft. radius of the explosion. The caliche is sorted out and transported to the leaching works. The caliche is extracted with water, and the solution is recrystallized so as to separate the sodium nitrate from the accompanying impurities—sodium chloride, sodium and calcium sulphates, sodium iodate, sodium perchlorate, insoluble matter, etc. Commercial Chili saltpetre contains from 95 to 98 per cent. of sodium nitrate. The sodium iodate which accumulates in the mother liquid is used for the manufacture of iodine. Nearly $2\frac{1}{2}$ million tons of "soda nitre" were exported from Chili in 1911.

The conversion of sodium nitrate into potassium nitrate.—The Chili saltpetre is dissolved in about $1\frac{1}{2}$ times its weight of boiling water, and a solution of sylvine—potassium chloride, from the Stassfurt deposits—in three times its weight of water, is poured into the sodium nitrate solution. Sodium chloride at once separates as a fine granular precipitate: $\text{KCl} + \text{NaNO}_3 \rightleftharpoons \text{NaCl} + \text{KNO}_3$. The crystals are removed, and the mother liquid is evaporated to about half its original volume, and the sodium chloride again removed. The sodium chloride is washed with water to recover some of the potassium nitrate removed with the crystals, and the washings used for dissolving more of the raw material. The mother liquid is further concentrated by evaporation. Crude nitre crystallizes from the solution. This is purified by recrystallization from boiling water, which is stirred while cooling so that the nitre crystals may be small and granular—"nitre meal." The crystals are dried and packed in sacks for transport. The potash nitre still holds about half per cent. of sodium chloride. The by-product—pickling salt—is preferred to ordinary salt for pickling meat, probably because of the quantity of nitre it still contains.

§ 2. Nitre Plantations.

The earth is one of Nature's vast laboratories in which microscopic organisms perform incredible experiments.—ANON.

The origin of the nitre beds is not known. It is generally agreed that the nitrogen is of organic origin—animal or vegetable. Since immense deposits of guano have been found on some of the islands off the coast of Peru—*e.g.* the Chincha Islands—it has been suggested that the nitrogen is derived from the guano. If so, it is not clear where the phosphates have gone, since there is practically no calcium phosphate in the nitre beds. Of course, the soluble nitrates may have been leached from decayed guano in some other locality and deposited in their present form. But the problem of the origin of the deposits has not been satisfactorily solved.

When organic matter decays, say, in the soil of cattle yards and stables, ammonia and ammonium compounds are produced by the action of certain bacteria. If the soil be fairly dry, but not too dry, a white scum appears on the exposed surface. The scum is made up of small crystals of potassium nitrate; and, after a time, if the soil be extracted with water, and strained, the liquid, on evaporation, furnishes yellowish-brown crystals of crude nitre. The crude nitre can be purified by re-solution and crystallization. The white efflorescence sometimes seen on the walls of stables, etc. is sometimes, not always, due to the formation of nitrates in this way.

The first stage in the decomposition of the organic matter is due to the action of certain bacteria. Ammonia and ammonium compounds are formed along with other gases which produce the characteristic odour of putrefaction. The next stage in the process of decomposition is due to the action of a special bacterium—the nitrous ferment—which converts the ammonia into nitrous acid: $2\text{NH}_3 + 3\text{O}_2 = 2\text{H}_2\text{O} + 2\text{HNO}_2$. Another bacterium—the nitric ferment—transforms the nitrous into nitric acid: $2\text{HNO}_2 + \text{O}_2 = 2\text{HNO}_3$. The two last-named varieties were isolated by Winogradsky in 1891. By the agency of these three types of bacteria, the soil is constantly receiving fresh supplies of nitrates necessary for the growth of plants, and derived from the decomposition of the organic matter present in the soil.¹ The free acids are not really present in the soil because the alkalies or alkaline earths present interact with the acids producing the nitrates and nitrites. It is owing to these reactions that water, contaminated by drainage from surface soil, contains nitrates.

During the Napoleonic wars, France had great difficulty in procuring sufficient nitre for the manufacture of gunpowder. This led to the construction of nitre beds in various parts of the country. But when the French ports were thrown open, after these wars, the manufacture of nitre, in France, was abandoned because it could be imported more cheaply from India. The process is still used in a few localities—*e.g.* Sweden. Soil rich in humus, dung, or animal offal is piled into heaps with the *débris* from buildings, or with lime, or wood ashes. The heaps are protected from rain by sheds. A system of gutters or pipes may also

¹ Of course, as indicated shortly, nitrates occur in rain-water derived from the direct oxidation of nitrogen by electric discharges—lightning, etc.

distribute the liquid excretions of animals over the top of the heap. The piled mass is called a nitre plantation. Before long a white film of nitre "grows" on the windward face of the pile. This is scraped off regularly, and leached as indicated above. If lime be present, the resulting calcium nitrate is converted into potassium nitrate by the addition of wood ashes—potassium carbonate: $\text{Ca}(\text{NO}_3)_2 + \text{K}_2\text{CO}_3 = 2\text{KNO}_3 + \text{CaCO}_3$. The nitre is then purified by recrystallization.

In the hot dry countries of the East, India, Persia, Arabia, etc., particularly in the neighbourhood of villages where urine and other organic matters find their way into the soil owing to imperfect systems of "sewage disposal," the process of nitrification goes on rapidly. The soil is extracted with water every few years, and the nitrates, chiefly calcium and potassium nitrate, are extracted as indicated above. The product from the soil in the Valley of the Ganges (Bengal) is called *Bengal saltpetre*, it is principally potassium nitrate. If calcium nitrate be present, potassium carbonate is added to the aqueous extract from the soil so that calcium carbonate may be precipitated, and potassium nitrate remain in solution.

§ 3. The Nitrogen Cycle.

Nitrogen is indispensable to life. Neither animals nor ordinary plants possess the faculty of taking it directly from the air. Here certain families of micro-organisms come to the rescue. Their obscure but highly efficacious life's work consists in preparing nitrogen for assimilation by plants, which, in their turn, get it into proper shape for the supply of animal necessities. . . . The chain of evidence is nowhere broken; its links, though distinct, are inseparable. At the very bottom we meet an army of invisible creatures endowed, for the benefit of others, with special facilities. Unsuccoured by them, plants must starve as surely as animals would starve in a world destitute of plants. So the highest standeth not without the lowest.—ANON.

All living matter and the waste products of animals contain considerable quantities of combined nitrogen. It is a necessary constituent for the growth of living organisms. During the decay of organic matter through the agency of bacteria, part of the nitrogen finds its way back to the atmosphere, and part passes directly into the soil to be absorbed by plants. Animals cannot assimilate free nitrogen, and they are accordingly dependent upon the plants for their supply. Plants in light transform inorganic compounds into complex organic products surcharged with energy and which are necessary for the life of higher animals. Plants thus form a permanent link between the inorganic and the animal kingdoms. Plants cannot usually obtain their nitrogen direct from the atmosphere. Most plants get their nitrogen from the soil where it is present in the form of nitrates, ammonium salts, or other complex compounds. The organic matter in the soil is attacked by bacteria of various kinds, and part is converted into nitrates and part into free nitrogen. A certain amount is brought back from the atmosphere, during a rain storm, where it has been oxidized into ammonium nitrate by electric discharges—this more particularly in the polar regions. These supplies of available nitrogen, however, do not suffice to maintain the fertility of cultivated soils. It is therefore necessary to make good the constant draining of the available nitrogen by the cultivated plants. This is done by allowing nitrogenous

organic matter—manures—to decay on the soil, or to add a mixture—fertilizer—containing available nitrogen.

In 1853 G. Ville noticed that while most plants reduce the available nitrogen in soils, some plants, principally the leguminosæ—peas, beans, clover, lupins, etc.—enrich the soil so that more available nitrogen is present after a crop has been removed than before. In 1886 H. Hell showed that certain leguminosæ appear to live in a kind of partnership—symbiosis (from the Greek *συν* (syn), with; *βίωσις* (biosis), living)—with certain bacteria—e.g. the *bacterium radicola*. The bacteria appear to live as guests in nodules on the rootlets of their host, and probably also in the neighbouring soil. The nodules on the rootlets of a *Phaseolus* (bean) are illustrated in Fig. 199. The symbiotic bacteria convert the nitrogen of the atmosphere into a form available as food for the plant on which they live.

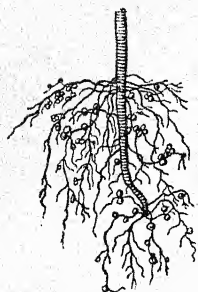
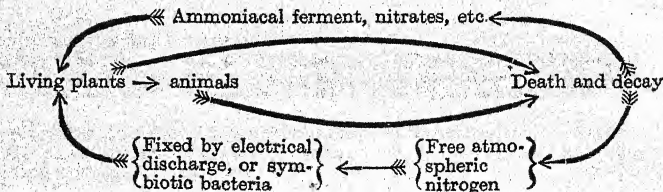


FIG. 199.—Nodules on Root of Bean Plant.

The processes concerned in the circulation of nitrogen in nature involve an eternal round of life and death, growth and decay; these can be summarized in the scheme:



The idea has been expressed in a more romantic way. To-day a nitrogen atom may be throbbing in the cells of the meadow grass; to-morrow it may be pulsating through the tissues of a living animal. The nitrogen atom afterwards may rise from decaying animal refuse, and stream to the upper regions of the atmosphere where it may be yoked with oxygen in a flash of lightning and return as plant food to the soil in a torrent of rain; or it may be directly absorbed from the atmosphere by the soil, and there rendered available for plant food by the action of symbiotic bacteria. Thus each nitrogen atom has doubtless undergone a never-ceasing cycle of changes through countless æons of time.

§ 4. The Fixation of Atmospheric Nitrogen.

The fixation of nitrogen is vital to the progress of civilized humanity, and unless we can class it among the certainties to come, the great Caucasian race will cease to be foremost in the world, and will be squeezed out of existence by the races to whom wheaten bread is not the staff of life.—W. CROOKES.

The fertility of cultivated fields and gardens is dependent upon the amount of combined nitrogen added as manure or fertilizer. Of course, cultivated plants require fertilizers containing other elements—particularly phosphorus, and potassium—but nitrogen is the most expensive.

The development of agriculture is largely dependent upon the cheap production of available nitrogen, and the cultivation of wheat, the staple food of the European races, is largely dependent on nitrogenous manures. The present supplies of nitrogen are mainly derived from (1) coal beds (gasworks, *vide* p. 873); (2) Chilian deposits; (3) nitre plantations and animal excrementa; (4) atmospheric nitrogen—*vide supra*; and (5) artificial methods of combining nitrogen with oxygen or hydrogen. According to a Report of a Commission of the Chilian Government in 1909, their nitre beds contained at least 250 million tons of caliche, containing over 15 per cent. of nitre, yet the consumption is so great—between two and three million tons per annum in 1912—that it was predicted the beds would be exhausted in less than a century. Most of the sodium nitrate is employed as a fertilizer for wheat, etc. Since the comparatively small store available promises soon to be depleted, it obviously became necessary to exploit other means of supplying farmers with the fertilizers they require. The nitre plantations do not give a large enough yield. F. Nobbe and L. Hiltner, in 1896, sold cultures of the "nitrogen-fixing" bacteria under the name "nitragen" for inoculating the soil. The results have been only fair for certain crops when peptones and glucose were added to the water in which the nitrifying bacteria are distributed for spreading on the soil.

There are at present three good methods for the conversion, or "fixation" of atmospheric nitrogen in a form available for plant food. (1) By heating a mixture of carbon with an alkali or alkaline earth base, or by heating calcium carbide in dry nitrogen whereby it is converted into calcium cyanamide (*q.v.*); (2) the direct synthesis of ammonia from its elements (*q.v.*); and (3) the direct oxidation of atmospheric nitrogen and absorption of the resulting oxides in water or alkaline solutions. It might be added that (4) the bacterial process (p. 616); and (5) the action of water on nitrides (p. 649) are not so promising.

Combustion process for oxidizing atmospheric nitrogen.—Near the end of the 18th century de Saussure showed that in all combustions a little atmospheric nitrogen is entrapped in the oxidation, and nitrogen oxides are accordingly formed. R. Bunsen also noticed the production of nitrogen oxides in his process for the analysis of gases, and M. Berthelot determined the amount of nitric acid formed when combustions are conducted in his bomb calorimeter with oxygen contaminated with nitrogen. The mechanism of the oxidation of nitrogen probably resembles that which occurs in the electric arc, *vide infra*. The excess of oxygen always present results in the formation of nitric acid. In Hoeüsser's process for making nitric acid, ordinary coal gas or coke-oven gas is detonated with an excess of air or of air mixed with oxygen in a bomb under pressure (5 atmospheres). The gases are rapidly cooled immediately after the explosion. In the continuous apparatus of Hoeüsser, fifteen explosions are produced per minute. In O. Bender's process the natural gas (CH_4 , 99.0 per cent.; H_2 , 0.4; N_2 , 0.6) of the gas-fields in Transylvania is burnt in burners of the Bunsen type under pressure.

Electric process for oxidizing atmospheric nitrogen.—J. Priestley (1775) first noticed that an acid is formed when electric sparks are sent through the air, but he seems to have thought that the acidity was due to carbonic acid. H. Cavendish (1785) proved that the product of the action is nitric

acid; while M. Berthelot showed that nitric oxide, NO , is an intermediate product. In Cavendish's experiment the air was confined over mercury in the bend of a Λ -shaped tube. The open ends of the tube dipped under mercury contained in separate glasses (Fig. 200). A series of sparks was sent from an electrical machine through the air confined in the tube. After the action, the gas turned blue litmus red, gave a turbidity with lime

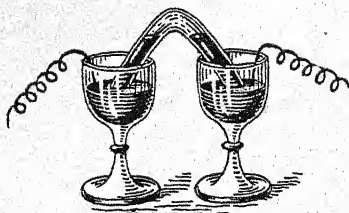


Fig. 200.—Cavendish's Experiment.

water, was absorbed by potassium hydroxide, etc. In fine the product of the action of potassium hydroxide upon the air after sparking was nitre. Z. Lefebvre obtained a patent in 1859 for making nitric acid by passing electric sparks through air and absorbing the gases in alkali bases; but the process does not appear to have been very successful. W. Crookes (1892) showed that air can be burned to nitric and nitrous acids in a power-

ful electric arc. With the ordinary "arc flame" the hot space is relatively small and the conditions for rapid cooling are bad. With intermittent sparking or arcing, the conditions for rapid cooling are better. The arc is spread out by a magnetic field, and accordingly the air in the arc is very hot; but just outside the arc the air is comparatively cold, so that the nitrogen oxides formed in the arc are rapidly cooled just outside the arc. Hence, when air is heated in the electric arc in order to oxidize the nitrogen, short thick arcs are to be avoided. Siemens and Halske (1902) burnt the nitrogen by passing air through a chamber containing an electric arc spread

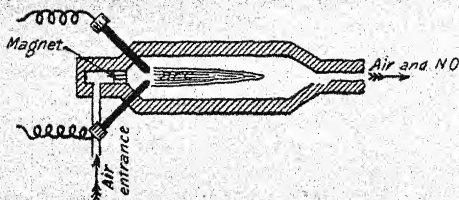


Fig. 201.—Siemens and Halske's Experiment.

over as great a surface as possible by means of an electro-magnet. Their apparatus is shown diagrammatically in Fig. 201.

The conditions of the reaction have been studied by W. Nernst (1906) and F. Haber (1907). Nitrogen is so stable that instead of exhibiting a

tendency to oxidize at ordinary temperatures, the nitrogen oxides tend to break down into their elements. As just indicated, nitric oxide is formed by the direct union of nitrogen and oxygen at high temperatures: $\text{N}_2 + \text{O}_2 = 2\text{NO}$. The reaction is endothermal. The higher the temperature, the greater amount of nitric oxide formed when the system is in equilibrium. Thus, with a mixture of equal volumes of nitrogen and oxygen:

Temperature . . .	1811°	2033°	2195°	3000°	3200°
Nitric oxide . . .	0.37	0.64	0.997	4.5	5.00 per cent.

The reaction is very rapid. A few seconds suffice for the system to assume equilibrium. Hence if the reaction is not to be reversed, owing

to the dissociation of nitric oxide, the system, after heating, must be cooled very rapidly, as was the case with hydrogen peroxide and ozone. Experiment shows that the cooling of the gases to about 700° suffices to make the back action (dissociation of nitric oxide) negligibly small, and the nitric oxide below 620° oxidizes to nitrogen peroxide, NO_2 .

There are several successful schemes for applying these facts commercially to the fixation of atmospheric nitrogen. Birkeland and Eyde's may be taken as typical. It is used in Norway. Other schemes are in use in the United States, Germany, Norway, Italy, etc. K. Birkeland and S. Eyde (1905) produce a high-voltage arc between two electrodes consisting of copper tubes through which a current of cold water is continually flowing. In order to spread the flame over as great an area as possible, an electromagnet is placed at right angles to the electrodes so that the terminals lie between the poles of the magnet. The effect of the magnet is to spread the flame on one electrode until the current is reversed; a new flame then starts on the opposite electrode. The current alternates every $\frac{1}{50}$ second, and the extreme rapidity of making and breaking the arc gives it the appearance of an intensely hot disc of flame—"electric sun"—6 feet in diameter—Fig. 202, which presents a large surface to the air. The "electric sun" is enclosed in a special brick-lined furnace with a steel casing, and air is driven past both sides of the disc of flame. The gases are pumped off. The "disc arc" (1) offers a very large surface of contact to the air. This means that a relatively large percentage of the nitrogen will be oxidized to nitric oxide with a minimum consumption of current; and (2) it allows the products of the reaction to be cooled rapidly so as to reduce the amount of nitric oxide dissociated during the cooling. The gases containing between 1 and $1\frac{1}{2}$ per cent. of nitric oxide, at a temperature of about 200° , enter the oxidation chamber where the nitric oxide combines directly with oxygen to form nitrogen peroxide, $2\text{NO} + \text{O}_2 = 2\text{NO}_2$. The gases are then passed through a series of five absorption towers where they meet water, and milk of lime. The absorbed nitrogen oxides form calcium nitrate; the solution is evaporated, and sold as fertilizer. The *Norwegian saltpetre* formed in Birkeland and Eyde's process is almost anhydrous calcium nitrate. It contains nearly 13 per cent. of available nitrogen. In illustration of the growth of this industry, 115 tons were produced in Norway in 1905, and 9422 tons valued at £72,590 were produced in 1909; and 110,000 to 120,000 tons in 1913; and in 1920, the capacity of installed works was 363,800 tons per annum.

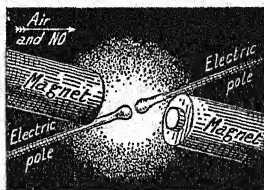


FIG. 202.—Electric Sun
(Diagrammatic).

§ 5. Nitric Acid—Preparation.

Molecular weight, $\text{HNO}_3 = 63.02$. Melting point, -47° ; boiling point, 86° at 760 mm. pressure. Specific gravity at 0° , 1.56.

History.—Nitric acid was probably not known to the ancient Egyptians. Geber says that he made it by distilling copperas with saltpetre and alum; and J. R. Glauber (1650) made it by distilling a mixture of nitre and

sulphuric acid. A. L. Lavoisier (1776) proved that nitric acid is a compound of oxygen, and H. Cavendish (1784-85) demonstrated that it is formed by sparking nitrogen with moist oxygen. J. L. Gay-Lussac (1816) found the ratio of hydrogen : oxygen : nitrogen corresponded with $\text{H}_2\text{O}.\text{N}_2\text{O}_5$.

Preparation of nitric acid.—When potassium or sodium nitrate is mixed with dilute sulphuric acid, in the cold, no obtrusive sign of chemical action occurs, although it can be proved that a reversible change has taken place so that the sodium is distributed between the sulphuric and nitric acids, and the same mixture is ultimately obtained whether the original constituents are sulphuric acid and sodium nitrate, or nitric acid and sodium sulphate—the latter reaction being endothermal, the former exothermal. If a mixture of concentrated sulphuric acid and sodium nitrate be heated to about 130° , nitric acid, HNO_3 , is volatilized. The reaction is represented : $\text{NaNO}_3 + \text{H}_2\text{SO}_4 \rightleftharpoons \text{NaHSO}_4 + \text{HNO}_3$. The two salts sodium nitrate and sodium hydrogen sulphate are not volatile ; the nitric acid, HNO_3 , boils at 86° ; and the sulphuric acid at 330° . Hence on warming to about 100° , much of the nitric acid is volatilized and the state of equilibrium of the solution is disturbed ; in consequence, the sodium nitrate is all decomposed. If the oily liquid in the retort be cooled, a little unchanged sodium nitrate may be crystallized along with crystals of $\text{NaHSO}_4.\text{H}_2\text{SO}_4$.

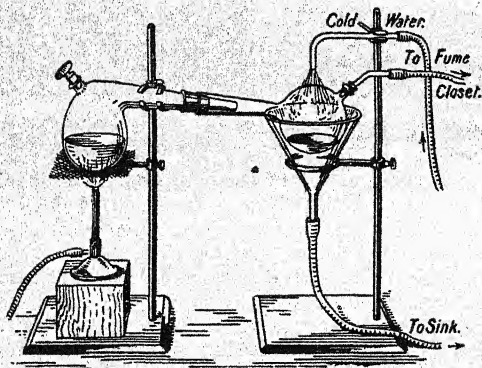


FIG. 203.—Preparation of Nitric Acid.

If the temperature of the retort was 120° , the crystals in the retort are almost wholly sodium hydrosulphate, NaHSO_4 ; and at still higher temperatures normal sodium sulphate is formed and hence less sulphuric acid is needed for a given yield of nitric acid : $2\text{NaNO}_3 + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{HNO}_3$. But appreciable quantities of the nitric acid are decomposed at the higher temperature.

For the manufacture of nitric acid from ammonia, *vide infra*.

The acid can be prepared in the laboratory by means of the apparatus illustrated in Fig. 203. This explains itself. All rubber and cork stoppers and connections must be avoided because the acid rapidly attacks organic matter. The retort is charged with the sodium nitrate and sulphuric acid, and heated. Brownish-red fumes appear and the distillate is more or less coloured brown because of the solution of the red-coloured gas in the distillate. The brownish-red is a product of the decomposition of the nitric acid by heat. On a manufacturing scale, the sodium nitrate and sulphuric acid are heated in cast-iron retorts ; the vapour is condensed in earthenware pipes cooled by water, and collected in earthenware jars. The last jar is connected with a tower filled with coke down which a stream

of water trickles. The object is to recover the nitrogen peroxide produced by the decomposition of the nitric acid. The retort has an exit pipe from which the sodium sulphate can be run when the action is over. To reduce the amount of nitrogen peroxide formed during the decomposition of the nitric acid by heat, the stills are often worked under a reduced pressure so that the acid may come off at as low a temperature as possible.

Purification of nitric acid.—The nitric acid so obtained contains some chlorine and iodine derived from the chlorides and iodides associated with the nitre. Some sodium sulphate, sulphuric acid, and iron are also carried over into the receiver. Nitrogen peroxide is also present as indicated above. To purify the acid, it is distilled in glass retorts, and the first fraction which come over is put on one side as crude acid containing volatile chlorine compounds. When the distillate gives no precipitate with a dilute solution of silver nitrate, the receiver is changed, and the greater part of the nitric acid is distilled off. The residue in the retort contains the sulphates, iodine and iron.

The acid can be redistilled from concentrated sulphuric acid to remove all the water; and the nitrogen peroxide can be removed by passing a current of dry carbon dioxide through the warm acid until it is colourless. The result is anhydrous nitric acid, HNO_3 .

Fuming nitric acid is brown in colour; it is nitric acid with a considerable amount of nitrogen peroxide in solution. It can be made by distilling nitric acid with a little starch. The starch reduces some of the nitric acid to nitrogen peroxide which is absorbed by the distillate. The fuming nitric acid of commerce contains 85 to 95 per cent. of HNO_3 .

Uses.—Nitric acid is one of the common acids. It is used as a solvent for metals, for etching designs upon copper; for the manufacture of nitrates used for photography [AgNO_3]; pyrotechny [$\text{Ba}(\text{NO}_3)_2$, $\text{Sr}(\text{NO}_3)_2$, etc.]; calico printing [$\text{Pb}(\text{NO}_3)_2$, etc.]. It is also used in the manufacture of dyestuffs, explosives, sulphuric acid, etc., and in metallurgy, etc. A carboy of the acid (sp. gr. 1.42) holds about 140 lbs. Commercial yellow fuming acid has about 47 per cent. HNO_3 , and the red fuming acid, 99.5 per cent. of HNO_3 . Crude commercial acid of specific gravity 1.35 has about 55 per cent. HNO_3 by weight, and commercially pure acid can be purchased with from 60 to 98 per cent. of nitric acid by weight; the acid with a specific gravity 1.4 has about 65 per cent. of HNO_3 . The chief impurities in commercial nitric acid are sulphuric and hydrochloric acids, iodine, nitrogen oxides, and arsenic. The commercial acid sells at about $3\frac{1}{2}d.$ per lb., and the pure at $4\frac{1}{2}d.$ per lb.

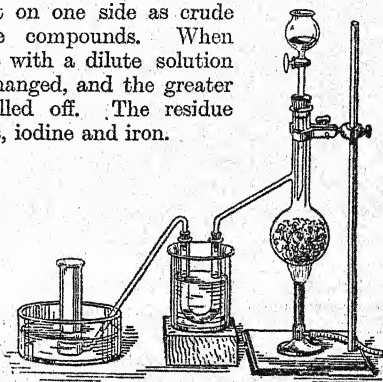


FIG. 204.—Decomposition of Nitric Acid by Heat.

§ 6. The Properties of Nitric Acid.

Nitric acid is a colourless mobile liquid which fumes strongly in air. It has a peculiar smell. The pure acid is hygroscopic and rapidly absorbs moisture from the air. It mixes in all proportions with water; and, as in the case of sulphuric acid, nitric acid contracts when mixed with water, and the mixture rises in temperature. The great contraction is said to correspond with a mixture of acid and water in the molecular proportions $3\text{H}_2\text{O} + \text{HNO}_3$. The pure acid boils at 86° , and freezes to a white solid, melting at -47° . An aqueous solution containing 68 per cent. of nitric acid boils at 120.5° , more concentrated solutions, and also more dilute solutions, boil at lower temperatures. A more dilute solution loses water on boiling, and a more concentrated solution loses acid on boiling until 68 per cent. of nitric acid of constant boiling point distils unchanged. This is the concentrated nitric acid of commerce. The specific gravity of the constant boiling acid at 15° is 1.414.

Decomposition of nitric acid.—A certain amount of acid is decomposed during the distillation: $4\text{HNO}_3 = 4\text{NO}_2 + 2\text{H}_2\text{O} + \text{O}_2$. This decomposition commences below 68° . At higher temperatures the decomposition is very marked. For instance, if the concentrated acid be allowed to fall drop by drop into a quartz flask containing fragments of calcined pumice-stone heated in an apparatus fitted as shown in Fig. 206, the red vapours evolved can be condensed in a freezing mixture to a brownish liquid (NO_2), and a colourless gas—oxygen—collects in the gas jar. Nitric acid vapour also decomposes when exposed to the light, nitrogen peroxide is formed which is absorbed by the acid giving it a yellow colour.

Hydrates of nitric acid.—F. W. Küster and R. Kremann's (1904) freezing curve for mixtures of nitric acid and water, Fig. 205, shows the existence of two hydrates: $\text{HNO}_3 \cdot 3\text{H}_2\text{O}$ and $\text{HNO}_3 \cdot \text{H}_2\text{O}$, see p. 525. The

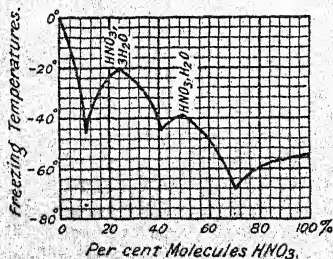


FIG. 205.—Freezing Curves of Nitric Acid and Water.

first-named hydrate separates in transparent crystals, melting at -18.2° ; and the latter in opaque crystals melting at -38° . Another hydrate, $\text{HNO}_3 \cdot 2\text{H}_2\text{O}$, is reported, but it does not appear to be stable under the conditions of this experiment.

With the notation employed in discussing the periodic acids, the acid $\text{N}(\text{OH})_5$ should be called orthonitric acid; $\text{NO}(\text{OH})_2$, mesonitric acid; and $\text{NO}_2(\text{OH})$, metanitric acid. The last alone is known with any degree of probability. The nitrates of inorganic chemistry are generally meta-salts.

The hydrate $\text{HNO}_3 \cdot 2\text{H}_2\text{O}$ has been referred to orthonitric acid, and $\text{HNO}_3 \cdot \text{H}_2\text{O}$ to mesonitric acid, but these compounds are not polybasic acids, and it is therefore doubtful if the hydrates of nitric acid can be interpreted in this way. H. Erdmann claims to have isolated orthonitric acid by conducting dry air through nitric acid at -15° , and under reduced pressure; the acid in question, $\text{HNO}_3 \cdot 2\text{H}_2\text{O}$, or $\text{N}(\text{OH})_5$, crystallizes in long needles which melt at -35° , and are quite stable below -15° .

Action on organic compounds.—Nitric acid is extremely corrosive, and when brought in contact with the skin causes painful sores. The dilute acid stains the skin yellow or brown, probably owing to the formation of xanthoproteic acid. If strong nitric acid be poured on sawdust, the mass often bursts into flame. If a dish of fuming nitric acid be placed in a basin in the bottom of a glass cylinder, and a little turpentine be added from a pipette, the turpentine will burst into flame. Glowing charcoal continues to burn when plunged into the acid. With organic compounds, one or more hydrogen atoms are replaced by an equivalent NO_2 radicle, and water is formed. For instance, cellulose or cotton, $\text{C}_{12}\text{H}_{20}\text{O}_{10}$, becomes nitrocellulose or gun cotton, $\text{C}_{12}\text{H}_{14}\text{O}_{10}(\text{NO}_2)_6$; glycerol, $\text{C}_3\text{H}_5(\text{OH})_3$, becomes nitroglycerol, $\text{C}_3\text{H}_5(\text{ONO}_2)_3$, which is the active agent in dynamite; and phenol, $\text{C}_6\text{H}_5\text{OH}$, becomes picric acid, $\text{C}_6\text{H}_2(\text{OH})(\text{NO}_2)_3$, which is the active agent in melinite.

Oxidizing action of nitric acid.—In consequence of the great proportion of oxygen in nitric acid, and in consequence of the ease with which it is decomposed, we should expect nitric acid to be a strong oxidizing agent. It is so. Sulphur is oxidized to sulphuric acid; e.g. $\text{S} + 2\text{HNO}_3 = \text{H}_2\text{SO}_4 + 2\text{NO}$; and also phosphorus to phosphoric acid, iodine to iodic acid, arsenious oxide to arsenic oxide, antimony to antimonious oxide, ferrous to ferric salts, etc. Many metallic sulphides—iron, lead, etc.—are oxidized by nitric acid to sulphates: $2\text{FeS}_2 + 10\text{HNO}_3 = \text{Fe}_2(\text{SO}_4)_3 + \text{H}_2\text{SO}_4 + 10\text{NO} + 4\text{H}_2\text{O}$.

Aqua regia.—A mixture of hydrochloric and nitric acids—say, in the proportions 3 or 4 volumes of the former per one volume of the latter—will dissolve the metals gold and platinum. Hence the alchemists term for this mixture, *aqua regia*—the kingly water—because it dissolves the very king of metals, gold. The mixture becomes coloured at about 10° , and it evolves gas at about 30° ; the action is quite vigorous about 90° , and at 100° a mixture of nitric and hydrochloric acids distils over. The action of *aqua regia* on gold is supposed to be due to the oxidation of the hydrochloric acid by the nitric acid whereby chlorine is formed, and this latter agent attacks the gold or platinum. Goldschmidt's equation is: $\text{HNO}_3 + 3\text{HCl} = 2\text{H}_2\text{O} + \text{NOCl} + \text{Cl}_2$, corresponding with the fact that some nitrosyl chloride— NOCl —is formed at the same time. The dissolution of gold in *aqua regia* is represented: $\text{Au} + \text{HNO}_3 + 4\text{HCl} = \text{NO} + 2\text{H}_2\text{O} + \text{HAuCl}_4$, and a mixture of nitric and hydrochloric acids in these proportions is considered to be the most economical for the purpose. The greater oxidizing properties of *aqua regia* over nitric acid are generally attributed to the presence of the free chlorine. The action on metallic sulphides, MS , is represented: $3\text{MS} + 6\text{HCl} + 2\text{HNO}_3 = 4\text{H}_2\text{O} + 2\text{NO} + 3\text{MCl}_2 + 3\text{S}$. This equation is imperfect since there is no doubt that the reaction is much more complex.

The 12th-century Latin versions of the writings of the Arabian chemist Geber describe the preparation of *aqua regia* by distilling a mixture of nitre, salt, and sulphuric acid. During the process, much chlorine must have been given off, and the properties of the gas are such that it must have made its presence known. J. B. van Helmont, J. R. Glauber, and R. Boyle must also have had the gas under observation, but they did not make much of it. J. R. Glauber said that when spirit of salt acts on certain metallic oxides, a gas the colour of fire is given off, and it attacks all metals.

Action on the metals.—The purest acid does not attack carbonates, nor does it dissolve mercury, copper, silver, cadmium, but sodium takes fire in the acid. Most metals and the carbonates are vigorously attacked by ordinary nitric acid, and for this reason the old alchemists called it *aqua fortis*—the strong water; or *aqua dissolutiva*—the dissolving water. The acid does not attack the so-called noble metals, gold, platinum, rhodium, and iridium. Tin with the dilute acid gives tin nitrate, $\text{Sn}(\text{NO}_3)_2$; but with the concentrated acid it forms nitrogen peroxide, NO_2 , and probably $\text{Sn}(\text{NO}_3)_4$, which is immediately decomposed by the heat of the reaction, so that stannic oxide, SnO_2 , alone appears as the solid product of the action: $\text{Sn}(\text{NO}_3)_4 = \text{SnO}_2 + 4\text{NO}_2 + \text{O}_2$. With mercury, an excess of the acid produces mercuric nitrate, $\text{Hg}(\text{NO}_3)_2$; and with an excess of mercury, mercurous nitrate, HgNO_3 . The action of nitric acid on the metals generally is somewhat complex, because the main reaction is complicated by side or concurrent, and by consecutive reactions. These again depend not only upon the particular metal under consideration, but also on the purity and concentration of the acid, the temperature, and the concentration of the products of the reaction accumulating in the solution. For instance, when dilute nitric acid acts upon copper, copper nitrate and a colourless gas, nitric oxide, are formed: $3\text{Cu} + 8\text{HNO}_3 = 3\text{Cu}(\text{NO}_3)_2 + 4\text{H}_2\text{O} + 2\text{NO}$; but as the copper nitrate accumulates in the solution, nitrous oxide and even nitrogen may be found in appreciable quantity among the products of the reaction; again, with concentrated nitric acid, nitrogen peroxide is the main gaseous product of the reaction: $\text{Cu} + 4\text{HNO}_3 = \text{Cu}(\text{NO}_3)_2 + 2\text{H}_2\text{O} + 2\text{NO}_2$. Similarly with zinc, dilute nitric acid forms nitrous oxide: $4\text{Zn} + 10\text{HNO}_3 = 4\text{Zn}(\text{NO}_3)_2 + 5\text{H}_2\text{O} + \text{N}_2\text{O}$; but with a more concentrated acid, ammonia may be formed: $4\text{Zn} + 9\text{HNO}_3 = 4\text{Zn}(\text{NO}_3)_2 + 3\text{H}_2\text{O} + \text{NH}_3$. The ammonia, of course, reacts with some of the nitric acid to form ammonium nitrate. Iron filings or a copper-zinc couple in the presence of a little sulphuric acid reduces dilute nitric acid to ammonia.

According to C. Montemartini (1892), at about 5°, the amount of ammonia increases with increasing concentration of acid and attains a maximum with 40 to 45 per cent. HNO_3 , after that the amount of ammonia decreases rapidly and very little is produced in a 60 per cent. acid. The curve, Fig. 63, shows the effect of the concentration of nitric acid upon the amount of ammonia produced per gram of zinc dissolved. At higher temperatures the maximum is attained with acids of less concentrations, thus, at 85°, the maximum ammonia formation occurs with a 9 per cent. acid. Secondary actions interfere with the regularity of the curve with the more concentrated acid. Some hyponitrous acid is probably formed. The quantity of nitrous oxide and nitrogen formed is small, the maximum amount of the former occurs with a 40 per cent. acid. The maximum velocity of the reaction is obtained with a 25 per cent. acid, with a more concentrated acid the speed of dissolution diminishes. Magnesium behaves like zinc, manganese too behaves somewhat similarly. With cadmium, iron, nickel, and cobalt, tin, antimony and molybdenum, ammonia is formed with very dilute acid; nitric oxide is produced, so also is nitrous oxide. If no ammonia is formed, neither nitrogen nor nitrous oxide are found in the gaseous products. Comparatively large quantities of nitrogen are produced with cobalt. No ammonia is found in the products of the acid

on copper, lead, bismuth, aluminium, mercury, and silver, but nitrous acid, nitric oxide, nitrogen peroxide are formed. Excluding lead, ammonia was only detected with those metals which can react with water at relatively low temperatures, and even lead can probably react with water at ordinary temperatures since it is soon oxidized. Fig. 206, although

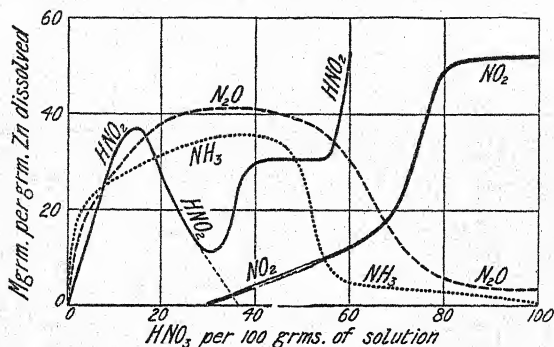
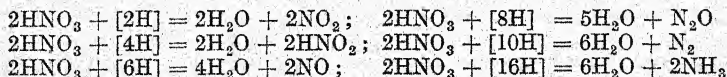


FIG. 206.—Products of the Action of Nitric Acid of Various Concentrations on Zinc (at 5°).

it does not indicate all the products of the reaction, illustrates in a striking manner the complexity of the reactions which occur when a metal dissolves in nitric acid.

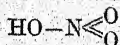
The mechanism of the reaction.—Some consider that the first product of the reaction is a nitrate of the metal and nascent hydrogen: $\text{Cu} + 2\text{HNO}_3 = \text{Cu}(\text{NO}_3)_2 + [\text{2H}]$. The nascent hydrogen is then supposed to reduce the nitric acid to nitrogen peroxide and nitrous acid. With some metals, the reduction of the nitric acid proceeds much further, say through the stages: $\text{NO}_2 \rightarrow \text{HNO}_2 \rightarrow \text{NO} \rightarrow \text{N}_2\text{O} \rightarrow \text{N}_2 \rightarrow \text{NH}_2\text{OH} \rightarrow \text{NH}_3$. Ammonia thus represents the limiting stage in the reduction of nitrogen compounds in the same sense that hydrogen sulphide represents the limiting stage in the reduction of sulphur compounds. Free hydrogen is seldom evolved because it is so rapidly oxidized by the nitric acid. However, free hydrogen is said to have been obtained during the action of nitric acid on manganese, and on magnesium. The reducing actions indicated above can be represented symbolically:



Some believe that the acid first oxidizes the metal to the oxide, e.g. $3\text{Cu} + 2\text{HNO}_3 = 3\text{CuO} + \text{H}_2\text{O} + 2\text{NO}$; and that the oxide then dissolves in the acid to form the nitrate: $\text{CuO} + 2\text{HNO}_3 = \text{Cu}(\text{NO}_3)_2 + \text{H}_2\text{O}$. When differences of opinion can reasonably be entertained, it follows that our knowledge of the facts is incomplete, and more experimental work is needed. He who doubts should inquire. The metals copper, silver, mercury, and bismuth, be it observed, have no perceptible action on cold dilute sulphuric and hydrochloric acids, and accordingly it is not likely that they will reduce nitric acid by the action of nascent hydrogen. In harmony with

what precedes, V. H. Veley (1890) proved that these metals have no action on cold dilute nitric acid unless a trace of nitrous acid or a lower nitrogen oxide is present. Nitrous acid may be present in the nitric acid as an impurity; it may be formed by the incipient decomposition of nitric acid when it is warmed; or it may be formed in the acid by electrolysis produced by local currents of electricity set up by impurities in the metal ("local action"). Once the action has started, the evolution of nitric oxide, and the formation of nitrate proceed quickly. According to Veley, therefore, the dissolution of this metal in nitric acid proceeds: $\text{Cu} + 3\text{HNO}_3 = \text{Cu}(\text{NO}_3)_2 + \text{HNO}_2 + \text{H}_2\text{O}$, is the resultant of a series of consecutive reactions: $\text{Cu} + 4\text{HNO}_2 = \text{Cu}(\text{NO}_2)_2 + 2\text{H}_2\text{O} + 2\text{NO}$; the copper nitrite is decomposed by the excess of nitric acid forming nitrous acid, and the nitric oxide formed reduces nitric to nitrous acid: $\text{Cu}(\text{NO}_2)_2 + 2\text{HNO}_3 = \text{Cu}(\text{NO}_3)_2 + 2\text{HNO}_2$; the small trace of nitrous acid thus acts as a catalytic agent; nitrous acid is continuously produced, and continuously decomposed: $3\text{HNO}_2 = \text{HNO}_3 + 2\text{NO} + \text{H}_2\text{O}$. The rate of solution of copper, silver, mercury, and bismuth in nitric acid is less when the solution is agitated than when it is at rest, since, in the latter case nitrous acid accumulates in the neighbourhood of the metal and accelerates the action.

Nitric anhydride; nitrogen pentoxide.—Pure nitrogen pentoxide is prepared by the fractional distillation of a mixture of phosphorus pentoxide and well-cooled pure nitric acid. The phosphorus pentoxide removes the elements of water from the nitric acid: $\text{P}_2\text{O}_5 + 2\text{HNO}_3 = 2\text{HPO}_3 + \text{N}_2\text{O}_5$. The mixture is distilled in a current of ozonized oxygen, and the vapours are passed over phosphorus pentoxide. The ozone oxidizes the lower oxides of nitrogen. The actual boiling and melting points cannot be fixed on account of sublimation. This substance is also made by the action of dry chlorine on silver nitrate in a U-tube between 60° and 95° . The same oxide is made when ozone acts on nitrogen peroxide. The nitrogen pentoxide reacts with water producing nitric acid: $\text{N}_2\text{O}_5 + \text{H}_2\text{O} = 2\text{HNO}_3$. Hence the crystalline mass is to be regarded as nitric anhydride. H. St. C. Deville (1849) analyzed this compound by passing the vapour over hot copper. The copper retained the oxygen, and the nitrogen passed on: $\text{N}_2\text{O}_5 + 5\text{Cu} = 5\text{CuO} + \text{N}_2$. The results of the analysis correspond with: nitrogen 25.65 per cent., oxygen 74.35 per cent., that is, with the formula $(\text{N}_2\text{O}_5)_n$. The molecular weight has not been determined. Hence we take the simplest— N_2O_5 . The structure of nitric acid is nearly always represented by



corresponding with quinquivalent nitrogen. The evidence is of a general kind: (1) the acid probably contains a hydroxyl group; and (2) the nitrogen is probably quinquivalent. The formula agrees with the ready decomposition of the pentoxide into nitrogen peroxide, NO_2 , and oxygen. The formula of the pentoxide is best represented as indicated above.

The freezing points of solutions of nitric acid, HNO_3 , in nitrobenzene, ethylene bromide, acetic acid, and chloroacetic acid, are but slightly higher than corresponds with the normal molecule, HNO_3 ; in boiling ethereal solutions, the molecular weight is normal. The density of the vapour of nitric acid at 86° is 59.2 ($\text{H}_2 = 2$), and at 256° , 36.0. The former is

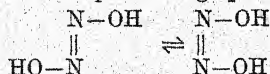
supposed to correspond with a 9.5 per cent. dissociation of the vapour HNO_3 ; and the latter with a 100 per cent. dissociation: $4\text{HNO}_3 \rightleftharpoons 4\text{NO}_2 + 2\text{H}_2\text{O} + \text{O}_2$.

§ 7. Hyponitrous Acid and the Hyponitrites.

God ever worketh by geometrical plan.—PLUTARCH.

Reduction of nitrites and nitrates.—The reduction of nitrites and nitrates to ammonia readily occurs in alkaline solutions. For example, if an excess of sodium or potassium hydroxide be added to nitric acid, and pure zinc, aluminium, or iron be added to the alkaline solution, the salt is reduced to ammonia. When a solution of potassium or sodium nitrite or nitrate is treated with a solution of metallic sodium in mercury—sodium amalgam—the hydrogen liberated reduces the nitrite or nitrate, forming a new salt—sodium hyponitrite: $2\text{NaNO}_3 + [\text{SH}] = 4\text{H}_2\text{O} + \text{Na}_2\text{N}_2\text{O}_2$. The alkaline solution can then be neutralized by acetic acid; and when silver nitrate is added, a yellow precipitate of silver hyponitrite, $\text{Ag}_2\text{N}_2\text{O}_2$, is formed. By treating silver hyponitrite, suspended in dry ether, with a solution of hydrogen chloride in dry ether, and evaporating the clear solution, white deliquescent crystalline plates are obtained. This is hyponitrous acid. The solid is very unstable and is liable to explode even below 0° . Barium hyponitrite, $\text{Ba}(\text{NO})_2$, can be made by the electrolysis of a mixture of equal parts of sodium nitrite and barium acetate in 12 times its weight of water. When the product is treated with dilute sulphuric acid, an aqueous solution of hyponitrous acid is formed. The acid can be extracted from its aqueous solution by ether, and the ether can be removed by evaporation. The aqueous solution soon decomposes into water and nitrous oxide: $\text{H}_2\text{N}_2\text{O}_2 = \text{H}_2\text{O} + \text{N}_2\text{O}$. The same substance, hyponitrous acid, is produced when hydroxylamine acts upon nitrous acid: $\text{NH}_2\text{OH} + \text{HNO}_2 = \text{H}_2\text{N}_2\text{O}_2 + \text{H}_2\text{O}$.

Constitution.—Hyponitrous acid is dibasic, and the formula $\text{H}_2\text{N}_2\text{O}_2$ has been established by freezing-point methods. Its constitution is best represented by $\text{HO}-\text{N}=\text{N}-\text{OH}$. It is conceivable that the molecule may have two different structures represented graphically by:



This means that with our present system of representing the composition of compounds, the hydroxyl groups in hyponitrous acid may be disposed either on the same side of the molecule—the *syn* type; or on opposite sides of the molecule—the *anti* type. Evidence of the existence of both forms has been obtained in organic chemistry. It will be observed that this form of isomerism may or may not exhibit a phenomenon resembling desmotropism, or tautomerism.

A striking example of this phenomenon was discovered by L. Pasteur in 1860. Tartaric acid—

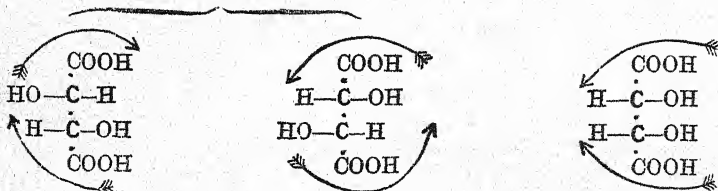


—exists in three forms; all have the same molecular weight, and the only apparent difference in their properties is connected with their behaviour towards polarized light. The most satisfactory method of representing the

constitution of the three tartaric acids in harmony with the facts, and with the atomic hypothesis, is as follows :

TRANS FORMS.

Cis FORM.



Dextro-tartaric acid.

Lævo-tartaric acid.

Inactive Tartaric acid.

There is one cis structure, and two modifications of the trans structure—the dextro- and lævo-acids.

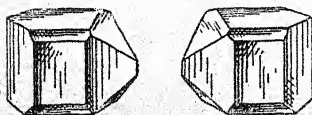


Fig. 207.—Crystals of Dextro- and Lævo-tartaric Acids.

The cis structure is inactive towards polarized light, while one of the trans forms deflects the ray of polarized light to the right (dextro-tartaric acid); and the other to the left (lævo-tartaric acid). There is no difference in the kind or number of atoms or radicles in the molecules of the different varieties. Crystals of the dextro- and lævo-acids or their salts are similar, but enantiomorphous—

from the Greek *ἐναντίος* (enantios), opposite; *μορφή* (morphe), shape—meaning that the two crystals are related to one another as object and image, or as right- and left-handed screws—Fig. 207. It is therefore inferred that the radicles in the two varieties are so disposed that their molecules rotate or spin in opposite directions; and that the third inactive variety (mesotartaric acid) is a kind of hybrid of the other two in that the tendencies to rotate in opposite directions neutralize one another. This is illustrated by the arrows in the above structural formulæ.

Instead of representing the disposition of the radicles in the molecules on a plane surface, it is maintained that the structural formulæ ought to be represented in three dimensions, as is usually done in organic chemistry, where this subject has been highly specialized. Hence the terms stereoisomerism—from the Greek *στερεός* (stereos), solid—or geometrical or physical isomerism. Many of the concepts which have been developed in organic chemistry have been extended into inorganic chemistry. The “syn” and “anti” forms of hyponitrous acid are simple examples. Sometimes the carbon atom behaves as if its valency linkages were directed from the centre towards the four apices of a regular tetrahedron. The real shape of the atom of carbon is, of course, unknown. Similar attempts have been made to deal with the atom of nitrogen and some other elements, but in these cases there is not the same agreement among chemists as is the case with carbon. See “Werner’s theory of valency.”

§ 8. Nitrates.

Nitric acid is a strong acid, it colours blue litmus red; and behaves as a monobasic acid with metallic oxides, hydroxides, and carbonates, forming a series of salts called nitrates. For instance, with copper oxide :

copper nitrate is formed: $\text{CuO} + 2\text{HNO}_3 = \text{Cu}(\text{NO}_3)_2 + \text{H}_2\text{O}$, etc. Similarly we have potassium nitrate, KNO_3 ; calcium nitrate, $\text{Ca}(\text{NO}_3)_2$; bismuth nitrate, $\text{Bi}(\text{NO}_3)_3$, etc.

The nitrates are usually readily soluble in water and form well-defined crystals. The nitrates are decomposed when heated. Lead nitrate, and the nitrates of the heavy metals generally, form an oxide of the metal and give off a mixture of oxygen and nitrogen oxides: $2\text{Cu}(\text{NO}_3)_2 = 2\text{CuO} + 4\text{NO}_2 + \text{O}_2$. The oxide of the heavy metal may not be stable at the temperature of the decomposition of the nitrate. With lead nitrate: $2\text{Pb}(\text{NO}_3)_2 = 2\text{PbO} + \text{O}_2 + 4\text{NO}_2$. The reaction is reversible at 357° , for if the nitrate be heated in an evacuated sealed tube, red fumes appear which vanish again when the tube is slowly cooled. Mercuric nitrate similarly forms mercuric oxide. The nitrates of the alkalis form nitrites and oxygen: $2\text{KNO}_3 = 2\text{KNO}_2 + \text{O}_2$. Hence the nitrates are often used as oxidizing agents.

Silver nitrate, AgNO_3 .—Silver nitrate is usually formed by the action of nitric acid on the metal. The aqueous solution crystallizes in colourless rhombic plates, 100 grams of water at 0° dissolve 122 grams of the salt; at 50° , 455 grams; and at 100° , 952 grams of the salt. The aqueous solution is blackened, probably by reduction to silver, by contact with organic matter, and it is accordingly used as marking ink for linen, etc. The crystals melt at 209° , and solidify to a fibrous crystalline mass called "lunar caustic." The salt decomposes at 450° , forming silver nitrite, AgNO_2 ; and at higher temperatures decomposes into metallic silver and nitrogen oxides. Silver nitrate absorbs ammonia gas with some avidity, forming silver triaminonitrate, $\text{AgNO}_3 \cdot 3\text{NH}_3$; the aqueous solution when saturated with ammonia deposits rhombic prisms of silver diamminonitrate, $\text{AgNO}_3 \cdot 2\text{NH}_3$, or $\text{Ag}(\text{NH}_3)_2\text{NO}_3$, or the salt is precipitated by adding alcohol or ether to the ammoniacal solution of silver nitrate. The same salt is precipitated when a solution of silver nitrate in benzonitrile is treated with ammonia. The salt is dissociated in aqueous solution and the silver is precipitated as chloride in the usual manner.

Composition and basicity of nitric acid.—To show how the basicity and composition of the acid can be determined: One gram of pure silver was dissolved in an excess of nitric acid, and on evaporation to dryness, a weight 1.5748 gram of silver nitrate was obtained. This silver nitrate was heated in a current of carbon dioxide, and the vapours passed over hot metallic copper. The copper retained the oxygen. The gases were collected over a solution of potassium hydroxide which retained the carbon dioxide, not the nitrogen. The volume of nitrogen collected corresponded with 0.1307 gram. Assuming that the atomic weight of silver is 107.88; oxygen, 16; it follows that:

	Grams.
Weight of silver nitrate	1.5748
Weight of metallic silver	1.0000
Weight of "nitrate" radicle	0.5748
Weight of nitrogen	0.1307
Weight of oxygen	0.4441

Hence in silver nitrate, the $\text{N}:\text{O} = 0.1307 : 0.4441$ by weight; i.e. the numbers of the atoms are related as $\text{N}:\text{O} = 0.0093 : 0.028$; or, as 1 : 3. Hence for every atom of silver, there are three atoms of oxygen and one

atom of nitrogen— AgNO_3 . Assuming this composition for silver nitrate, it follows that the combining weight of nitrogen can be determined directly from the data given. No one has prepared two silver nitrates—acid and normal silver nitrates—and there is no evidence of dibasicity with nitric acid. Hence it is inferred that nitric acid is monobasic.

Copper nitrate, $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$.—An aqueous solution of copper nitrate is obtained by the typical methods for the preparation of salts—the action of nitric acid on the metal, oxide, hydroxide, or carbonate. The solution deposits deliquescent crystals of $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$, and when heated to about 60° , the crystals decompose, forming basic copper nitrate, $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{Cu}(\text{OH})_2$. Anhydrous copper nitrate, $\text{Cu}(\text{NO}_3)_2$, has not been made.

Mercurous nitrate, HgNO_3 , is deposited in colourless monoclinic crystals of $\text{HgNO}_3 \cdot \text{H}_2\text{O}$ from solutions of mercury in cold dilute nitric acid (mercury in excess). The effect on the freezing of aqueous solutions corresponds with the formula $(\text{HgNO}_3)_2$. The salt dissolves in water acidulated with nitric acid, but an excess of water decomposes the salt with the precipitation of a basic nitrate, $\text{HgOH} \cdot \text{HgNO}_3$. If this be boiled with water, it is converted into mercuric nitrate and mercury; if an excess of mercury be also present, the basic mercurous nitrate $3\text{HgNO}_3 \cdot 2\text{HgOH}$ is formed. Mercuric nitrate, $\text{Hg}(\text{NO}_3)_2$, is prepared by boiling mercury with an excess of nitric acid until the solution gives no precipitate with a little sodium chloride. If evaporated over sulphuric acid, deliquescent crystals of $2\text{Hg}(\text{NO}_3)_2 \cdot \text{H}_2\text{O}$ are formed. If the mother liquid be boiled, a compound $\text{Hg}(\text{NO}_3)_2 \cdot \text{HgO} \cdot 2\text{H}_2\text{O}$ is precipitated, and if this precipitate, or mercuric nitrate, be treated with an excess of cold water, $\text{Hg}(\text{NO}_3)_2 \cdot 2\text{HgO} \cdot \text{H}_2\text{O}$ is precipitated as basic mercuric nitrate. Thus, like mercurous nitrate, mercuric nitrate has a great tendency to form basic salts.

Lead nitrate, $\text{Pb}(\text{NO}_3)_2$.—Lead nitrate is formed by dissolving the metal, the oxide, carbonate, etc., in nitric acid. The salt is deposited in regular octahedral crystals isomorphous with barium nitrate. When heated lead nitrate decomposes as indicated above. The aqueous solution is faintly acid, and basic salts are said to be formed when the aqueous solution is boiled with lead monoxide— $\text{Pb}(\text{NO}_3)_2 \cdot \text{OH}$, and when lead nitrate is heated to, say, 300° in a closed tube. Other basic salts are obtained by adding ammonia to a solution of lead nitrate. The composition depends, as is so often the case, with the so-called basic salts, on the concentration and temperature of the solution. Quite a complex series of mixed products is obtained by the action of lead and of potassium nitrite on lead nitrate.

§ 9. Nitrous Acid and the Nitrites.

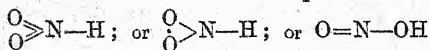
Sodium or potassium nitrates decompose when heated, forming corresponding nitrite, but when specially preparing these nitrites, it is best to mix the nitrate with a mild reducing agent like metallic lead or iron filings. Thus by heating at 470° a mixture of sodium nitrate and an excess of lead oxide: $3\text{PbO} + \text{NaNO}_3 \rightarrow \text{NaNO}_2 + \text{Pb}_3\text{O}_4$; or by heating a mixture of lead filings with an equal weight of sodium nitrate in a crucible: $3\text{Pb} + 4\text{NaNO}_3 = \text{Pb}_3\text{O}_4 + 4\text{NaNO}_2$. Digest the cold mass with water, filter, and evaporate the solution to a small bulk, when crystals of sodium nitrite separate on cooling. Most of the nitrites are very soluble in water, but silver nitrite, AgNO_2 , is not very soluble in

cold water. It is precipitated in crystalline needles when a solution of silver nitrate is added to a solution of sodium nitrite. If carefully purified, potassium and sodium nitrites have a neutral, not an alkaline, reaction.

Nitrous acid can be made by adding hydrochloric or sulphuric acid to a solution of the nitrite: $\text{NaNO}_2 + \text{HCl} = \text{NaCl} + \text{HNO}_2$. If hydrogen peroxide be added to a solution of ammonia, the latter is oxidized to nitrous acid: $\text{NH}_3 + 3\text{H}_2\text{O}_2 = \text{HNO}_2 + 4\text{H}_2\text{O}$. Since an excess of ammonia is usually present, ammonium nitrite is formed. When ammonium nitrite is heated, nitrogen is formed (p. 677). The aqueous solution of the acid gradually decomposes at ordinary temperatures: $3\text{HNO}_2 = \text{HNO}_3 + 2\text{NO} + \text{H}_2\text{O}$. The decomposition is much quicker when the solution is warmed. Nitrous acid is only known in dilute solutions.

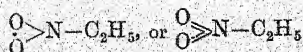
The acid acts both as an oxidizing agent and as a reducing agent. Thus, it *oxidizes* hydriodic acid or potassium iodide liberating iodine: $2\text{HI} + 2\text{HNO}_2 = 2\text{H}_2\text{O} + 2\text{NO} + \text{I}_2$. A solution of potassium permanganate, however, is rapidly decolorized in the presence of sulphuric acid with the formation of manganous sulphate. The permanganate is thus *reduced* by the nitrite solution.

Constitution of nitrous acid.—The formula HNO_2 is confirmed by electrical conductivity measurements, and by the effect of the acid on the freezing of water. There is some difference of opinion which of the formulæ:



represents nitrous acid. The difference between the first two can only be answered by evidence showing whether the nitrogen atom be ter- or quinquevalent. The latter formula is generally accepted on the assumptions: (1) The acid probably contains a hydroxyl group; and (2) the nitrogen atom is probably trivalent. As a matter of fact the evidence shows that both types of formulæ may be right under different conditions. The formation of ethyl nitrite, $\text{C}_2\text{H}_5\text{NO}_2$, by distillation from a mixture of ethyl alcohol, sodium nitrite, and sulphuric acid; and the subsequent hydrolysis of ethyl nitrite to ethyl alcohol and water may be taken to show that the ethyl radicle is attached to oxygen, not to nitrogen: $\text{O}=\text{N}-\text{OC}_2\text{H}_5 + \text{NaOH} = \text{C}_2\text{H}_5\text{OH} + \text{O}=\text{N}-\text{ONa}$. Ethyl nitrite is reduced by tin and hydrochloric acid to ammonia, alcohol, and some hydroxylamine: $\text{O}=\text{N}-\text{OC}_2\text{H}_5 + 3\text{H}_2 = \text{H}_2\text{O} + \text{NH}_3 + \text{C}_2\text{H}_5\text{OH}$.

Again, if silver or sodium nitrite be heated with ethyl iodide, $\text{C}_2\text{H}_5\text{I}$, in a sealed tube, the so-called nitroethane, $\text{C}_2\text{H}_5\text{NO}_2$, is produced. Nitroethane and ethyl nitrite are isomeric, the former boils at 113° , the latter at 17° . When reduced with tin and hydrochloric acid, nitroethane forms ethylamine, $\text{C}_2\text{H}_5\text{NH}_2$, a compound also produced by the action of ethyl iodide on ammonia. Nitroethane is not hydrolyzed by sodium hydroxide solution. These facts are taken to mean that in nitroethane, the nitrogen is attached directly to the ethyl group:



Both nitroethane and ethyl nitrite are derived from sodium nitrite, and it is therefore inferred that nitrous acid exhibits tautomerism:



Nitrocellulose and the other compounds indicated on p. 623 are in harmony with one of the first two formulae, because in these compounds the group NO_2 acts as a monad radicle. On the other hand, compounds like nitrosyl chloride, NOCl ; nitrosyl fluoride, NOF , etc., agree with the third formula, because in these compounds the OH radicle is replaced by monad chlorine, fluorine, etc. Ortho-nitrous acid may be regarded as the final oxidation product of ammonia, where NH_3 becomes $\text{N}(\text{OH})_3$. This orthonitrous acid is unknown even in combination. Ordinary nitrous acid is metanitrous acid, and the nitrites are meta-salts.

Identification of nitrites and nitrates.—(1) Dilute sulphuric acid decomposes all the nitrites in the cold. Nitric oxide is formed, which gives brownish-red fumes in the air: $3\text{HNO}_2 = \text{HNO}_3 + 2\text{NO} + \text{H}_2\text{O}$. Nitrates do not give the brown fumes, unless they be heated with concentrated sulphuric acid. (2) Potassium iodide is not decomposed by pure dilute nitric acid; whereas, with a solution of a nitrite, when acidified with dilute sulphuric or acetic acid, it becomes yellow or brown owing to the separation of iodine. If a little starch paste be present, the solution will be coloured blue. Many other oxidizing agents give the same coloration—ozone, hydrogen peroxide, chlorine, and ferric salts. Nitrates give the same coloration if a little zinc be added to the acidified solution owing to the reduction of the nitrates to nitrites. (3) A sensitive test for the nitrates is based upon the oxidation of ferrous salts by nitric acid: $6\text{FeSO}_4 + 3\text{H}_2\text{SO}_4 + 2\text{HNO}_3 = 3\text{Fe}_2(\text{SO}_4)_3 + 4\text{H}_2\text{O} + 2\text{NO}$; and by nitrous acid: $2\text{FeSO}_4 + \text{H}_2\text{SO}_4 + 2\text{HNO}_2 = \text{Fe}_2(\text{SO}_4)_3 + 2\text{H}_2\text{O} + 2\text{NO}$. In the cold, the nitric oxide forms a dark brown compound with the excess of ferrous salt. To apply the test, dissolve the nitrate in as little water as possible. Add a cold saturated solution of ferrous sulphate slightly acidified with sulphuric acid. Pour concentrated sulphuric acid down the side of the tube. If nitric acid be present, the zone of contact will be coloured dark brown. With nitrites, the concentrated sulphuric acid need not be added. (4) Nitrates may be distinguished from nitrites by the addition of a hot concentrated solution of sodium hyposulphite. The former remain unaltered, the latter evolve nitrogen.

§ 10. Nitrous Oxide, or Nitrogen Monoxide.

Molecular weight, $\text{N}_2\text{O} = 44.02$. Melting point, -102.7° ; boiling point, -89.8° ; critical temperature, $+36.5^\circ$. Vapour density ($\text{O}_2 = 32$), 44.02 ; (air = 1) 1.5299 . One litre under standard conditions weighs 1.9777 grams.

Preparation.—Nitrous oxide, as we have seen, is produced when nitric acid, specific gravity 1.1, reacts with zinc or tin. J. Priestley discovered the gas in 1772, and he made it by reducing nitric oxide with moist iron filings, or potassium sulphide, etc. The gas is also formed by heating hydroxylamine nitrite: $\text{NH}_2\text{OH} \cdot \text{HNO}_2 = 2\text{H}_2\text{O} + \text{N}_2\text{O}$. Nitrous oxide, however, is rarely prepared by these methods for experimental work.

One most convenient mode of preparation is to heat dry ammonium nitrate in a flask fitted with a delivery tube (Fig. 103). The salt melts at about 165° , and it begins to decompose at about 185° . The decomposition proceeds quite rapidly between 200° and 240° . Nitrous oxide and water are formed: $\text{NH}_4\text{NO}_3 = 2\text{H}_2\text{O} + \text{N}_2\text{O}$. At 240° the decomposition is very vigorous, and the mixture may even explode. There are also side

reactions, for the nitrate may decompose: $2\text{NH}_4\text{NO}_3 = 2\text{N}_2 + \text{O}_2 + 4\text{H}_2\text{O}$; and $2\text{NH}_4\text{NO}_3 = \text{N}_2 + 2\text{NO} + 4\text{H}_2\text{O}$. Each of the three reactions evolves heat. The maximum degradation occurs when nitrogen, oxygen, and water are the products of the reaction, and since nitrous oxide is an endothermal compound which itself readily breaks down into nitrogen and oxygen, the formation of nitrous oxide may be regarded as an intermediate product in the thermal decomposition of ammonium nitrate. A well-dried mixture of ammonium sulphate and sodium nitrate is often preferred to ammonium nitrate because the decomposition then proceeds quietly and more uniformly: $(\text{NH}_4)_2\text{SO}_4 + 2\text{NaNO}_3 = \text{Na}_2\text{SO}_4 + 4\text{H}_2\text{O} + 2\text{N}_2\text{O}$. When ammonium nitrate decomposes very rapidly, appreciable quantities of nitric oxide are formed. If required for special work, the gas must be purified from chlorine (formed from the chlorides contaminating the ammonium nitrate), by passing the gas through a solution of potassium hydroxide, and from nitric oxide, by passing the gas through a solution of ferrous sulphate.

Properties.—Nitrous oxide is a colourless gas with a faint smell. 100 c.c. of water, at 760 mm. pressure, dissolve 130 volumes of the gas at 0° ; and 67 volumes at 20° . Hence the gas is usually collected over hot water in order to lessen the loss due to its solubility in that liquid. The aqueous solution has a sweetish taste. The gas is about four times as soluble in alcohol as in water. The gas condenses to a colourless limpid liquid at 0° under a pressure of thirty atmospheres. M. Faraday liquefied the gas in 1823. Liquid nitrous oxide can be purchased in steel cylinders. The liquid boils at -89.3° , and freezes to a snow-like mass when allowed to evaporate. The cubic crystals melt at -102.7° . The liquid forms a crystalline hydrate: $\text{N}_2\text{O} \cdot 6\text{H}_2\text{O}$, with water. Both the liquid and solid produce painful blisters when dropped on the hand. If liquid nitrous oxide be mixed with carbon disulphide and placed in a vacuum, the temperature of the mixture falls to -140° .

Nitrous oxide resembles oxygen in its behaviour towards combustibles. A brightly glowing splinter bursts into flame when plunged into the gas. Ignited phosphorus, sulphur, etc., burn vividly in nitrous oxide gas. The burning of, say, sulphur in nitrous oxide gas with a greenish-blue flame is sometimes represented as: $2\text{N}_2\text{O} + \text{S} = \text{SO}_2 + 2\text{N}_2$, when, as a matter of fact, if the combustion be carried out in a large vessel, the walls become covered with crystals of nitrosulphonic anhydride and some of its decomposition products; some nitrogen peroxide is also formed. The reaction is therefore quite a complicated one. When J. Priestley discovered oxygen in 1774, he was already familiar with nitrous oxide, which he had discovered a couple of years previously. Hence in describing the brilliancy of the flame of a burning candle in oxygen, he said: "I got nothing like this remarkable appearance from any kind of air besides this particular modification of nitrous air." The burning body decomposes the nitrous oxide, unites with the oxygen, and leaves the nitrogen as a residue. If sulphur be but feebly burning, its flame is extinguished when plunged into nitrous oxide, probably because the temperature is not high enough to decompose the gas. It is therefore easy to mistake nitrous oxide for oxygen. One distinguishing test is to add a bubble of suspected gas to the nitric oxide; if red fumes are produced the gas is oxygen. Nitrous oxide does not give red fumes with nitric oxide. Another test depends on the far greater

solubility of nitrous oxide in alcohol than oxygen. Nitrous oxide is reduced by moist magnesium, zinc, cadmium, or iron, and some ammonia is formed at the same time.

When inhaled, nitrous oxide produces unconsciousness, and insensibility to pain. Hence it has long been used as an anæsthetic for small surgical operations, dentistry, etc. But owing to the unpleasant after effects sometimes produced, it is not used so much as formerly. If the inhalation be long continued, it may produce death; while if but small quantities are inhaled, it may produce a kind of hysteria, or intoxication. Hence the gas is sometimes called "laughing gas." As L. Edgworth remarked (1799), after breathing the gas, "I burst into a violent fit of laughter, and capered about the room without having the power of restraining myself": and H. Davy wrote (1794): after breathing sixteen quarts of the gas, "I danced about the laboratory as a madman." The effects of the gas are not quite the same on different people.

The formation of nitrous oxide is an endothermal reaction: $2\text{N} + \text{O} = \text{N}_2\text{O} - 18 \text{ Cals.}$ Nitrous oxide decomposes with an explosion if a fulminating cap be detonated in the gas. Nitrous oxide decomposes into its elements when heated, $2\text{N}_2\text{O} \rightarrow 2\text{N}_2 + \text{O}_2$. Two volumes of the gas furnish two volumes of nitrogen and one volume of oxygen—three volumes in all.

Composition of nitrous oxide.—The oxygen can be withdrawn from nitrous oxide, by sodium or potassium, or barium sulphide, in an apparatus like that indicated in Fig. 170. The equation cannot be given because of the uncertainty what particular oxide is formed. With potassium

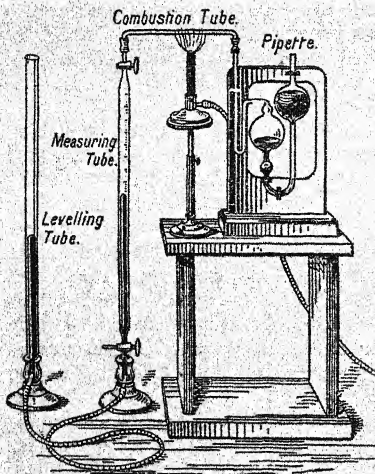


FIG. 208.—Composition of Nitrogen Oxides.

an oxide is formed which spontaneously absorbs oxygen when exposed to the air. The following is a better way of conducting the experiment. A hard glass or quartz tube about 3 mm. bore, and 10–12 cm. long is filled with metallic copper made by reducing the granulated oxide in a current of hydrogen. The copper is held in place by plugs of asbestos at each end. This tube is mounted between Hempel's burette (p. 115), and a two-bulbed gas pipette, Fig. 208. Both the burette and pipette are charged with mercury. The upper bulb of the pipette is empty when the lower bulb is full of mercury; the mercury extends to a mark on the gauge tube of the pipette. The burette contains a measured volume of nitrous oxide.

By opening the stopcock, and raising the levelling tube, a slow current of the gas is led over the red-hot metallic copper into the pipette. When the lower bulb of the pipette is nearly full of gas, return the gas to the burette by lowering the

levelling tube. When the gas is cold, read its volume when the mercury in the measuring and levelling tubes is at the same level. It will be found that although the gas has decomposed, and copper oxide¹ is formed, this has made no difference to the volume of the gas in the burette. This experiment proves that nitrous oxide contains its own volume of nitrogen. Or one molecule of nitrous oxide contains one molecule of nitrogen; that is, two atoms of nitrogen and the formula of nitrous oxide must therefore be N_2O_n where n is to be determined.

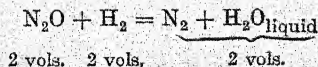
One gram-molecule of nitrous oxide weighs . . .	44.027 grams
One gram-molecule of nitrogen weighs . . .	28.020 "
Weight of oxygen in the molecule . . .	16.007 "

Assuming that the small fraction here represents an experimental error, the number 16.007 can represent one and only one atom of oxygen, since the atomic weight of oxygen is 16. This means that the formula for nitrous oxide is N_2O . The constitution of this apparently simple compound is by no means clear. The molecule is supposed to be constituted $\begin{smallmatrix} N \\ \vdots \\ N \end{smallmatrix} > O$

and the nitrogen atom is supposed to be trivalent, not univalent, as might be inferred by analogy with water $H.O.H$, or, by analogy with chlorine monoxide, Cl_2O , where chlorine is usually considered to be univalent. There are no grounds for assuming trivalent nitrogen in nitrous oxide other than the general statement that "nitrogen frequently acts as a trivalent element, and rarely, if ever, as a univalent element." J. Thiele supposed that the formula of nitrous oxide is best represented by the unsymmetrical formula $O=N\equiv N$, where one nitrogen atom is trivalent, and the other quinquivalent. Nitrous oxide appears to be the anhydride of hyponitrous acid: $N_2O + H_2O = H_2N_2O_2$, but nitrous oxide does not seem to react with water in this manner; although nitrous oxide is formed when concentrated sulphuric acid acts on hyponitrous acid: $H_2N_2O_2 + H_2SO_4 = H_2SO_4.H_2O + N_2O$. Nitrous oxide has also been analyzed gravimetrically by determining the increase in weight of the copper (Fig. 208); and also by heating electrically a weighed spiral of iron wire in the gas, and finding the increase in weight due to the absorption of oxygen, that is, to the formation of ferric oxide.

EXAMPLE.—1.1670 grams of nitrous oxide, gave 0.4242 gram of oxygen; and hence the gas contains 0.7428 gram of nitrogen. What is the formula of nitrous oxide? The gas contains 0.7428 gram of nitrogen per 0.4242 gram of oxygen. Divide these numbers by the corresponding atomic weights, as indicated on p. 78, and we get the atomic ratio $N : O = 2 : 1$. Hence the formula is $(N_2O)_n$. But the density of nitrous oxide is 44.0074. This corresponds with the molecule N_2O when the atomic weights of oxygen and nitrogen are respectively 16 and 14.01.

A mixture of equal volumes of nitrous oxide and hydrogen explodes when sparked, so that



The water condenses to a liquid, and hence its volume is negligibly small. The explosion is not so violent as with electrolytic gas, p. 69. Since

¹ Below 350° cuprous oxide Cu_2O , not CuO , is formed; above 350°, CuO is produced.

this method is used for determining the amount of oxygen admixed with nitrogen, the process may be described in a little more detail.

Analysis of compounds of nitrogen and oxygen.—Suppose that 30 c.c. of nitrous oxide be transferred to a Hempel's burette charged with mercury (p. 115), and then 40 c.c. of hydrogen be added. Connect the Hempel's burette with Hempel's explosion pipette, Fig. 209, by means of a piece of capillary tube and two pieces of thick-walled pressure tubing. The explosion pipette consists of a thick-walled glass bulb in which are fused two platinum wires with tips about 2 mm. apart. This explosion bulb is connected with another bulb by means of a glass tube and stopcock. The upper end of the explosion bulb ends in a U-shaped gauge tube. The explosion pipette is charged with mercury in such a way that the upper bulb

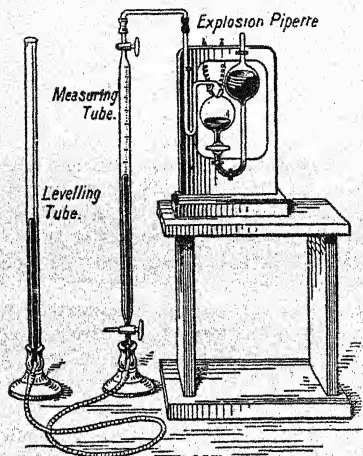


FIG. 209.—Analysis of Gases by Explosion.

is empty when the lower bulb is full of mercury, and the mercury extends to a mark on the gauge tube. Open the two stopcocks and raise the levelling tube of the burette so as to transfer the gas from the burette to the explosion pipette. Allow a little mercury to pass into the pipette from the burette to make sure that all the gas has been transferred to the pipette. Close both stopcocks, and pass a spark across the platinum terminals. After the explosion, open both stopcocks, and transfer the gas from the pipette to the burette by depressing the levelling tube until the level of the mercury in the burette is in its former position. The volume of the water formed during the explosion is negligibly small in comparison with the gases from which it is formed. Bring the

mercury to the same level in both tubes and read the volume of the gas in the burette. Suppose :

BEFORE EXPLOSION :	Volume of nitrous oxide	30 c.c.
	Volume of hydrogen added	40 c.c.
	Total volume	70 c.c.
AFTER EXPLOSION :	Total volume of mixture	40 c.c.
	Contraction	30 c.c.

If 30 c.c. of hydrogen have been converted into water ; 15 c.c. of oxygen must have been used for this reaction. Hence 15 c.c. of oxygen must have been taken from the nitrous oxide. Hence 10 c.c. of hydrogen were added in excess of what was actually required. Consequently, 30 c.c. of nitrogen remain mixed with 10 c.c. of hydrogen. Some oxides of nitrogen are formed during the explosion, and these dissolve in the water forming nitric and nitrous acids. This causes rather low results for nitrogen.

§ 11. Nitric Oxide, or Nitrogen Dioxide.

Molecular weight, $\text{NO} = 30.01$. Melting point, -167° ; boiling point, -150° ; critical temperature, -94° . Vapour density ($\text{H}_2 = 2$), 29.88 ; (air = 1) 1.039 .

History.—J. Priestley (1772) is generally regarded as the discoverer of nitric oxide; although J. Mayow (1669) made it by treating iron with nitric acid; and R. Boyle (1671) noted that it formed reddish fumes in contact with air. J. B. van Helmont (c. 1600) knew the gas, although his descriptions seem to confuse it with carbon dioxide, probably because he had one name—gas sylvestre—for a number of different gases.

Preparation.—Nitric oxide is prepared by the action of nitric acid, specific gravity 1.2, upon metallic copper or mercury. The copper turnings are placed in a two-necked Woulfe's bottle (Fig. 10) or a Kipp's apparatus (Fig. 13). The bottle is about one-fourth filled with water, and about the same volume of concentrated nitric acid is added. A rapid evolution of gas occurs. The gas should be collected as soon as possible because when the reaction has been in progress some time, particularly if the temperature rises during the reaction, nitrous oxide and nitrogen may appear with the nitric oxide, as indicated on p. 624. The results are better if a little sodium nitrite, say 2 per cent., be added to the mixture in the Woulfe's bottle.

A purer gas is obtained by reducing potassium nitrate with ferrous sulphate acidified with sulphuric acid, or ferrous chloride acidified with hydrochloric acid: $\text{KNO}_3 + 3\text{FeCl}_2 + 4\text{HCl} = 3\text{FeCl}_3 + \text{KCl} + 2\text{H}_2\text{O} + \text{NO}$. A mixture of 12 grams of potassium nitrate with 100 grams of ferrous sulphate is introduced into a flask, Fig. 105, and mixed with 100 c.c. of water and 60 c.c. of sulphuric acid. A steady stream of nitric oxide is evolved when the mixture is warmed. According to L. W. Winkler (1889), highly pure nitric oxide can be made by dropping a 50 per cent. solution of sulphuric acid upon a mixed solution of potassium iodide and sodium nitrite in the apparatus depicted Fig. 50; and F. Emich made a highly pure oxide by the joint action of sulphuric acid with about 2 per cent. of sodium nitrate and mercury. The gas was scrubbed in a tower containing potassium hydroxide.

Properties.—Nitric oxide is a colourless gas a little heavier than air. When brought in contact with air, it immediately combines with the oxygen, forming brownish-red fumes of nitrogen peroxide. No other gas gives red fumes when exposed to the atmosphere or to oxygen gas. Hence it is not possible to describe the smell, and the physiological action of this gas. If the two gases—nitric oxide and oxygen—be thoroughly dried, no combination occurs. This agrees with a generalization of M. Traube, "Dry substances do not unite directly with oxygen." If dry nitric oxide be passed into liquid oxygen or liquid air greenish flecks appear which F. Raschig attributed to the formation of nitrogen hexoxide. Analyses give results corresponding with the empirical formula NO_3 . Hence the reaction is represented $\text{NO} + \text{O}_2 = \text{NO}_3$. Nitrogen hexoxide decomposes spontaneously into a greyish-blue solid, nitrogen peroxide, N_2O_4 , and free oxygen, at temperatures a little above the boiling point of oxygen. E. Müller (1914) considers that the alleged hexoxide is really a mixture of N_2O_4 , N_2O_3 , and NO , because if a test-tube of liquid air be dipped into liquid nitrogen trioxide, a yellow crust of the latter congeals on the test-

tube; and if the tube be now dipped in liquid nitrogen peroxide, a blue layer of the latter is deposited on the former. The combination of the two coloured films, yellow and blue, gives a green identical with the alleged hexoxide.

At 0° and 760 mm. pressure, 100 volumes of water dissolve 7.3 volumes of the gas; and at 20°, 4.6 volumes. Nitric oxide slowly reacts with potassium hydroxide, forming nitrous oxide, nitrogen, and potassium nitrite. The reaction is very slow at ordinary temperatures, but at 125° the reaction is over in a couple of hours. There are probably two side reactions: $4\text{NO} + 2\text{KOH} = \text{N}_2\text{O} + 2\text{KNO}_2 + \text{H}_2\text{O}$; and $6\text{NO} + 4\text{KOH} = \text{N}_2 + 4\text{KNO}_3 + 2\text{H}_2\text{O}$. Nitric oxide dissolves in a solution of ferrous sulphate. A "compound" of nitric oxide and ferrous sulphate appears to be formed. This imparts a dark brown colour to the solution. The "compound" is decomposed when heated to about 60°, and nitric oxide is evolved. It is not very clear if a true compound is formed between the nitric oxide and the ferrous salt, because its composition seems to vary

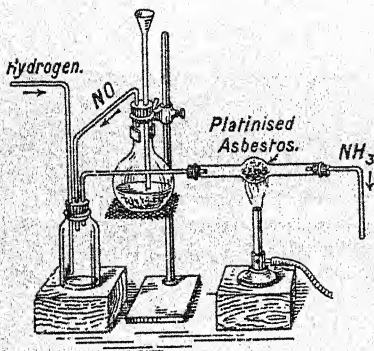


FIG. 210.—Reduction of Nitric Oxide to Ammonia.

of nitrous acid can be detected. $\text{KMnO}_4 + \text{NO} = \text{KNO}_3 + \text{MnO}_2$.

Nitric oxide is difficult to liquefy. At -93.5° , a pressure of 71.2 atmospheres is required to liquefy the gas. The liquid is colourless if air be excluded, otherwise the liquid may be tinted green or blue. The liquid boils at -150° , and the white solid melts at -167° .

Nitric oxide is not combustible, and it only supports combustion under special conditions, that is when the temperature is raised sufficiently to decompose the gas. The flame of feebly burning phosphorus is extinguished, but if the phosphorus be burning vigorously, combustion is continued in the gas. Burning sulphur is extinguished, but if the sulphur be boiling when it is plunged in the gas combustion sometimes continues. Potassium also burns in the gas, forming a mixture of nitrites and nitrates. Sodium can be heated in a tube containing the gas without the metal burning. A mixture of nitric oxide with hydrogen can be sparked without explosion, but if the mixture be passed through a hot tube, decomposition takes place: $2\text{H}_2 + 2\text{NO} = 2\text{H}_2\text{O} + \text{N}_2$. If the tube contains platinized

with the temperature of formation. Thus at 8° the composition corresponds with $3\text{FeSO}_4 \cdot 2\text{NO}$; from 8° to 25° , about $2\text{FeSO}_4 \cdot \text{NO}$; and above 25° , $5\text{FeSO}_4 \cdot \text{NO}$. With these facts, and the law of constant composition before us, we cannot very well say that ferrous sulphate forms a true chemical compound with nitric oxide, p. 570. Ferrous, cobaltous, nickelous, manganous, and chromous chlorides give similar results. Nitric oxide also dissolves in nitric acid forming a coloured solution ranging in tint from brown to yellow, to green, to blue. When passed through a solution of potassium permanganate, no formation The reaction is represented:

asbestos or finely divided nickel or copper (reduced from the oxides) ammonia gas is formed: $2\text{NO} + 5\text{H}_2 = 2\text{H}_2\text{O} + 2\text{NH}_3$. A slow current of nitric oxide is mixed with an excess of hydrogen in the apparatus illustrated in Fig. 210, which explains itself. The dry hydrogen comes from a Kipp's apparatus. The ammonia and the excess of hydrogen can be passed through a dilute solution of hydrochloric acid of sulphuric acid to form an ammonium salt; a red litmus paper held in the stream of gas will be coloured blue.

Decomposition.—When nitric oxide is exposed to the shock of a detonation from mercury fulminate, the gas decomposes into its elements. The experiment can be made by exploding a percussion cap by means of an electric spark in a tube of the gas. Nitric oxide, however, is one of the most stable of the nitrogen oxides. It does not decompose appreciably when heated until the temperature reaches 500° . The reaction $2\text{NO} \rightleftharpoons \text{O}_2 + \text{N}_2$ is reversible, as indicated on p. 618, so that at still higher temperatures, say, 3000° , over 4 per cent. of the mixture of nitrogen and oxygen will combine to form nitric oxide. This behaviour is characteristic of endothermal reactions, and $\text{N} + \text{O} = \text{NO} - 21.6 \text{ Cals.}$

Nitric oxide *reduces* potassium permanganate, iodic acid, silver oxide, nitric acid, etc., in the cold; and it *oxidizes* potassium and barium sulphides, etc., when warmed.

Composition.—The composition of nitric oxide may be established by the methods employed for nitrous oxide. The residual nitrogen occupies half the volume of the original gas. Hence one molecule of nitric oxide contains half a molecule of nitrogen; that is, one atom. The formula is therefore NO_n , where n has to be determined. The density of the gas is 29.88 ($\text{H}_2 = 2$). Hence,

One gram-molecule of nitric oxide weighs	29.88
One half gram-molecule of nitrogen weighs	<u>14.01</u>
Equivalent weight of oxygen	15.88

This number, 15.88, can represent one and only one atom of oxygen, atomic weight 16. The formula must, therefore, be NO . The vapour density agrees with this. Hence the term nitrogen dioxide for this gas appears to be a misnomer. The name nitrogen dioxide was given because nitric oxide contains twice as much oxygen for the same quantity of nitrogen as nitrogen monoxide— N_2O .

EXAMPLE.—R. W. Gray (1905) found that 0.6430 gram of nitric oxide, when passed over hot finely divided nickel, increased the weight of the nickel 0.3430 gram, and the liberated nitrogen, when condensed at the temperature of liquid air in coconut charcoal, furnished 0.3001 gram of nitrogen. What is the formula of nitric oxide taking P. A. Guye and C. Davila's (1905) number 1.3402 for the density of nitric oxide when air is unity, or 30.0124 when oxygen is 32? Divide 0.3430 by 16, and 0.3001 by 14.01. We get the atomic ratio 0.02144 : 0.02143, which is very nearly as 1 : 1. If the formula be NO , the vapour density is 30.01. This agrees with the observed result.

If nitrogen be a triad, and oxygen a dyad, there must be one free valency in the graphic formula, $-\text{N}=\text{O}$. The facts are, if oxygen is bivalent, nitrogen behaves as if it were also bivalent. The ready way in which nitric oxide unites with other elements by direct addition rather lends itself to the idea that nitric oxide has a free valency. Witness the direct combination of nitric oxide with oxygen at ordinary temperatures;

and the direct combination of nitric oxide with chlorine and fluorine forming respectively nitrosyl chloride, NOCl , and nitrosyl fluoride, NOF . The NO group is sometimes called nitrosyl, and compounds containing NO as a monad radicle, nitrosyls. In this sense, nitrogen peroxide may be regarded as nitrosyl oxide, NO_2O ; and nitrous acid, nitrosyl hydroxide, $\text{NO}-\text{OH}$.

Nitrosyl chloride, NOCl , when prepared by the slow union of gaseous chlorine and nitric oxide at ordinary temperatures, the speed of the reaction agrees with the assumption that the equation is $2\text{NO} + \text{Cl}_2 = 2\text{NOCl}$. The speed of the reaction is greatly stimulated in the presence of dry granular charcoal between 40° and 50° . Nitrosyl chloride can be prepared in many other ways. *E.g.* by the action of phosphorus pentachloride, PCl_5 , upon potassium nitrite: $\text{PCl}_5 + \text{KNO}_2 = \text{KCl} + \text{POCl}_3 + \text{NOCl}$; see also aqua regia, p. 623. Nitrosyl chloride is an orange yellow gas which condenses to an orange-coloured liquid at -8° ; and freezes to lemon-coloured crystals at -65° . It is decomposed by water into nitrous and hydrochloric acids: $\text{NOCl} + \text{H}_2\text{O} = \text{HNO}_2 + \text{HCl}$. It has no action on gold and platinum, but it attacks mercury, forming mercurous chloride, HgCl , and nitric oxide; it forms double compounds like $\text{SnCl}_4 \cdot \text{NOCl}$; $\text{FeCl}_3 \cdot \text{NOCl}$; etc. The density of the gas at 0° is 65.456 ($\text{O}_2 = 32$); this datum has been employed for checking the atomic weight of chlorine, given $\text{N} = 14$, and $\text{O} = 16$, it follows that $\text{Cl} = 35.46$. From measurements of the dissociation of nitrosyl chloride, and the union of nitric oxide and chlorine between 140° and 320° , M. Trautz (1913) considers that some other molecules are present—say, $(\text{NO})_2\text{Cl}$, $(\text{NO})_3\text{Cl}$, or NOCl_2 . Only between -72° and 40° does the brown gas consist wholly of molecules NOCl .

§ 12. Nitrogen Peroxide, or Nitrogen Tetroxide.

Molecular weight, $\text{NO}_2 = 46.01$; $\text{N}_2\text{O}_4 = 92.02$. Melting point, -9° ; boiling point, $+26^\circ$. Vapour density depends upon the temperature.

History.—As indicated in discussing nitric oxide, R. Boyle (1671) knew that nitric oxide formed red fumes in air, and these red fumes are mentioned in alchemical writings as *the blood of the salamander*. Since nitrogen peroxide is always formed during the preparation of nitric acid, nitrogen peroxide must have been recognized as a gas whenever nitric acid was made. J. L. Gay-Lussac (1816) first made its composition clear.

Preparation.—As indicated above, this gas is formed when one volume of oxygen is mixed with two volumes of nitric oxide: $2\text{NO} + \text{O}_2 = 2\text{NO}_2$. If the gas be led through a freezing mixture, the peroxide condenses either as a liquid or as a solid. Similarly, if the products obtained by heating a metallic nitrate (p. 629), say, lead nitrate, be similarly treated, this compound also condenses as a more or less impure liquid. Perhaps the best way of making nitrogen peroxide is to put, say, 200 grams of coarse lumps of arsenic trioxide, As_2O_3 , into a flask, and add 250 grams of nitric acid (specific gravity 1.4). Heat the mixture on a sand-bath, and lead the gases through a tower packed with glass wool, Fig. 211, and finally into a U-tube surrounded by a freezing mixture of ice and salt. A dark blue liquid is condensed. This is a mixture of nitrogen peroxide with

nitrogen trioxide, etc. When the evolution of gas has ceased, pass a current of air or oxygen through the condensed liquid while still in the freezing mixture until the liquid becomes yellowish-brown. In special cases, the yellowish-brown liquid can be further purified by mixing it with a large quantity of phosphorus pentoxide and strong nitric acid. Decant the nitrogen peroxide from the syrupy liquid, and redistil the mixture by warming its containing vessel in hot water, and passing the gases through the tower and condensing tube as in Fig. 211.

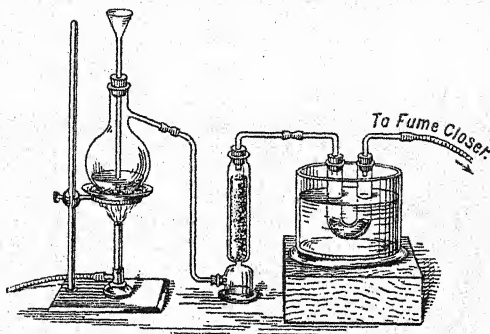


FIG. 211.—Preparation of Nitrogen Peroxide.

The bluish-grey powder formed when the alleged nitrogen hexoxide is washed with liquid nitrogen was supposed by F. Raschig to be an isomeric form of nitrogen tetroxide which he called nitrogen isotetroxide. This product is said to be stable at -188° , but when the temperature rises to -110° , it gives off oxygen, and forms a mixture of blue nitrogen trioxide and white crystals of nitrogen tetroxide, N_2O_4 .

Action of heat on nitrogen peroxide.—At low temperatures nitrogen peroxide forms colourless prismatic needle-like crystals. The crystals melt to a colourless liquid at -9° , but owing to "undercooling" a temperature in the vicinity of -30° is sometimes needed to solidify the liquid. As the temperature of the liquid rises, it begins to acquire a pale greenish-yellow tint, which becomes deeper and deeper, until at 10° , the liquid is distinctly yellow; at 15° , orange; and at 26° , the liquid boils and forms a reddish-brown vapour. The colour of the vapour becomes deeper and deeper until, at 40° , it is dark chocolate brown, and almost opaque; at 140° , the vapour is almost black. On cooling the vapour, the same changes occur in the reverse order. The effect seems to be connected with a change in the vapour density of the compound. The vapour density for N_2O_4 is 92.02; and for NO_2 , 46.01. The actual value of this constant passes from 84.7 at -72.6° and 115.4 mm. pressure, to 46.0 at 140° and 760 mm. pressure; and for intermediate temperatures:

Temperature	26.7°	60.2°	100.1°	135.0°	140°
Density ($H_2 = 2$)	76.6	69.0	48.6	46.2	46.0
NO_2 molecules	20.0	50.0	79.2	99.0	100.0 per cent.

Given the vapour density of the gas and the vapour density of each component, it is possible to calculate, by the subjoined method, the relative proportions of N_2O_4 and NO_2 molecules at any given temperature. Let x denote the fraction of the total volume of the gas present as NO_2 , then $1 - x$ will represent the fraction of the total volume present as N_2O_4 . The vapour density of NO_2 is 46, and of N_2O_4 , 92. Hence if the molecular weight of the gas be M , that is, the vapour density, hydrogen 2, we shall

have $M = 46x + (1 - x)92$. Hence unit volume of the gas will contain

$$x = \frac{92 - M}{46} \text{ volumes of NO}_2$$

EXAMPLE.—What proportion of nitrogen peroxide is present in 100 volumes of the gas when the vapour density is 76.6 ($H_2 = 2$)? Here, $M = 76.6$; and $x = 0.333$, that is, the gas contains 33.3 per cent. by volume of NO_2 .

In the preceding example, for every 0.333 volume of NO_2 there is 0.667 volume of N_2O_4 . Each molecule of the latter weighs twice as much as the former, so that the total weight of the gas must be $0.333 + 2 \times 0.666 = 1.667$ units when there is 0.333 unit of NO_2 molecules present. Hence there will be $0.333 \div 1.667 = 0.20$ unit of NO_2 by weight; or 20 per cent. of the total weight of the gas will be present as NO_2 molecules. The percentage proportion of NO_2 molecules in the gas at different temperatures

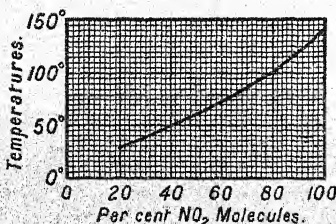
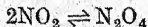


FIG. 212.—Dissociation Nitrogen Peroxide.

is indicated in Fig. 212. It is therefore probable: (1) Colourless nitrogen peroxide consists of N_2O_4 molecules, and coloured nitrogen peroxide of NO_2 molecules; (2) At low temperatures, the molecules are mainly N_2O_4 , and at high temperatures, NO_2 ; (3) The dissociation of N_2O_4 into NO_2 begins when the compound is in the liquid state; and (4) Nitrogen peroxide below 140° is a mixture of N_2O_4 and NO_2 molecules, and just above 140° , of NO_2 molecules alone.

By changing the pressure, keeping the temperature constant, similar changes occur. The less the pressure, the greater the number of NO_2 molecules. Again, when nitrogen peroxide is diluted with chloroform, the amount of N_2O_4 which dissociates increases with dilution as well as with rise of temperature. The freezing point of a solution of nitrogen peroxide in acetic acid corresponds with the molecule N_2O_4 .

The facts indicate that we are here dealing with two opposing reactions (p. 119) represented:



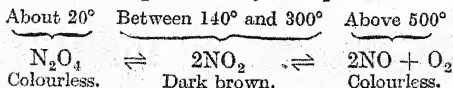
The velocity of the right to left transformation is proportional to the concentration of the N_2O_4 molecules (p. 117); that is, the rate of dissociation of the N_2O_4 molecules is equal to $k_1 C_1$, where C_1 denotes the concentration of the N_2O_4 molecules expressed in, say, gram-molecules per unit volume, and k_1 is the affinity constant, p. 119. Similarly, the rate of combination of the $NO_2 + NO_2$ molecules is proportional to the concentration of the reacting NO_2 molecules per unit volume, that is, to $kC \times C$, where C denotes the concentration of the reacting NO_2 molecules. For equilibrium, these two opposing reactions will be the same, and we shall have:

$$kC^2 = k_1 C_1; \text{ or } \frac{C^2}{C_1} = \frac{k_1}{k}; \text{ or } \frac{C^2}{C_1} = K$$

where K is the so-called equilibrium constant. The experimental results in all cases indicate that the equation just obtained: $C^2 = KC_1$ represents the condition of equilibrium of the gas with respect to the distribution

of the two different kinds of molecules at different temperatures and pressures. The numerical values of the affinity constants change with the temperature, but its value is constant for differences of pressure; and the observed results agree with the formula $C^2 = KC_1$.

At about 500° , an appreciable number of the dark brown molecules of NO_2 begin to dissociate into a colourless mixture of nitric oxide and oxygen: $2\text{NO}_2 = 2\text{NO} + \text{O}_2$ (cf. p. 639). The action of heat on nitrogen peroxide may therefore be represented by the equations:



Properties.—Nitrogen peroxide is a poisonous gas, and soon produces headache and sickness if but a little is present in the atmosphere.

Nitrogen peroxide is not combustible, and it extinguishes the flame of a taper. Phosphorus, sulphur, and carbon, if burning vigorously, may continue burning in the gas, but only when the temperature of combustion is sufficiently high to decompose the gas. Nitrogen peroxide is an energetic oxidizing agent. Phosphorus, carbon, potassium, mercury, copper, etc., when heated in the gas, are oxidized, while the gas is decomposed. The gas liberates iodine from potassium iodide; and it reduces permanganates, forming nitric acid. Nitrous oxide, indeed, is the only gaseous nitrogen oxide which does not decolorize aqueous solutions of potassium permanganate. Nitrogen peroxide resembles ozone and hydrogen peroxide in some of its reactions. The properties of these substances are compared side by side, see Table XXXV., p. 646.

Action of water.—Nitrogen peroxide is decomposed by water. At low temperature, a mixture of nitric and nitrous acids is formed: $\text{N}_2\text{O}_4 + \text{H}_2\text{O} = \text{HNO}_3 + \text{HNO}_2$, and with aqueous solutions of the alkalis, the corresponding salts are obtained. The solution of mixed nitrous and nitric acids is stable only when very dilute, otherwise the nitrous acid decomposes: $3\text{HNO}_2 = \text{HNO}_3 + 2\text{NO} + \text{H}_2\text{O}$. Summarizing these two equations: $3\text{NO}_2 + \text{H}_2\text{O} = 2\text{HNO}_3 + \text{NO}$. The reaction is reversible because when nitric oxide is passed into nitric acid, nitrogen peroxide is formed. When water is treated with an excess of nitrogen peroxide, the solution passes through a series of colour changes—blue, green, orange. This is due to the gradual solution of the nitrogen peroxide in the nitric acid formed as just indicated. If sufficient water be present, the solution finally becomes colourless. With warm water, the nitrous acid decomposes into nitric acid, and nitric oxide: $3\text{NO}_2 + \text{H}_2\text{O} \rightleftharpoons 2\text{HNO}_3 + \text{NO}$.

J. Priestley (1772) analyzed air by taking advantage of the fact that when nitric oxide is added to air, nitrogen peroxide is formed, and the red fumes are dissolved by potassium hydroxide. Hence, if nitric oxide be gradually added to a known volume of air, in presence of this alkali, the nitrogen peroxide is absorbed. The residual gas is nitrogen, all the oxygen is removed by the nitric oxide. This method is not often used because other processes are less troublesome. The formation of acid when nitrogen peroxide dissolves in water; and the formation of nitrogen peroxide by the action of oxygen on nitric oxide, may be illustrated by an apparatus fitted up as shown in Fig. 213. A little water is placed in the globe. The

globe is then filled with nitric oxide. The lower vessel is filled with water tinted with blue litmus. Oxygen is led slowly into the globe through the side tube. Red fumes of nitrogen peroxide are formed; this gas is absorbed by the water and the pressure is reduced. The coloured water rises from the dish into the globe to restore equilibrium. The blue litmus is coloured red by the acid formed in the globe. If everything is properly regulated, the globe will be nearly filled with water, and the first rush of water will appear as a miniature fountain inside the globe.

Nitroxyl.—Nitrogen peroxide unites directly with copper, cobalt, nickel, and iron, forming the so-called **nitro-metals**. For instance, finely divided copper obtained by reducing the oxide in a stream of hydrogen,

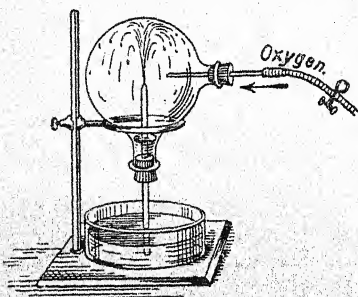
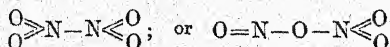


FIG. 213.—Nitrogen Peroxide Fountain.

absorbs about 1,000 times its volume of nitrogen peroxide, much heat is evolved, and brown solid copper nitroxyl, $\text{Cu}(\text{NO})_2$, is formed. The group NO_2 thus behaves as if it were a monad radicle; it is called **nitroxyl**; and compounds containing the NO_2 group are called **nitroxyls**. At 90° , copper nitroxyl is decomposed into metallic copper and nitrogen peroxide: $\text{Cu}(\text{NO})_2 \rightleftharpoons \text{Cu} + 2\text{NO}_2$.

Composition.—In gaseous nitrogen peroxide, NO_2 , we have to assume either that nitrogen is quadrivalent, $\text{O}=\text{N}=\text{O}$; or else that nitrogen is quinquivalent, $\text{O}=\text{N}=\text{O}$, with a

free valency. There is no free valency difficulty with the compound N_2O_4 , because two nitrogen atoms are here either quinquivalent; or one is quinquivalent, and the other trivalent:



The action of cold water on the peroxide forming nitrous and nitric acid seems to favour the second formula, *viz.* $\text{NO}-\text{O}-\text{NO}_2$. E. Divers argued in favour of $\text{O}=\text{N}-\text{O}$ for the formula of nitrogen dioxide, and $\text{O}:\text{N}:\text{O}-\text{O}:\text{N}:\text{O}$ for the tetroxide. Sometimes the nitrogen oxides are regarded as constituted like the water molecule with one or both the hydrogen atoms replaced by univalent radicles NO or NO_2 ; thus:



Water,
 H_2O .



Nitrous acid,
 HNO_2 .



Nitrogen tri-
oxide, N_2O_3 .



Nitrogen tetrox-
ide, N_2O_4 .



Nitrogen pentox-
ide, N_2O_5 .

Nitrogen peroxide appears to be a mixed anhydride of both nitric and nitrous acids. The composition of nitrogen peroxide can be established first by showing that it is formed by the union of equal volumes of oxygen and nitric oxide, by the apparatus indicated Fig. 208, and by vapour density determinations. There are several organic compounds known in which nitrogen acts as a quadrivalent element. The oxidation of a mixture of potassium hydrosulphite and nitrite furnishes a yellow salt with the

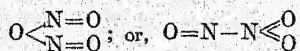
empirical composition $(\text{KSO}_3)_2\text{NO}$, and a violet salt of the same composition. The former is bimolecular, and the latter quadrimolecular. In the former, the nitrogen behaves as if it were quadrivalent $(\text{KSO}_3)_2=\text{N}=\text{O}$.

§ 13. Nitrogen Trioxide.

Molecular weight, $\text{N}_2\text{O}_3 = 76.02$. Melting point, -111° ; boiling point, 2° —with decomposition. The vapour density shows that the molecule decomposes.

When a mixture ¹ of nitric oxide and nitrogen peroxide is passed through a tube cooled to about -30° , a more or less impure form of nitrogen trioxide— N_2O_3 —condenses to a bluish liquid. As soon as the temperature rises, the liquid dissociates: $\text{N}_2\text{O}_3 = \text{NO} + \text{NO}_2$, nitric oxide escapes, and leaves a residual yellow liquid of nitrogen peroxide. The melting-point curve of mixtures of nitric oxide and nitrogen peroxide indicates the formation of a compound N_2O_3 melting at about -103° . If more nitrogen peroxide be present, the melting point falls to the eutectic temperature -112° . The trioxide is also produced by the direct union of solid nitrogen peroxide with nitric oxide; and by the action of oxygen on liquid nitric oxide; and as a green flocculent powder, resembling chromic hydroxide, when a series of electric discharges are passed through liquid air. The trioxide, not the peroxide, appears to be formed when nitric oxide unites with oxygen at temperatures below -100° . Nitrogen trioxide is also formed when electric discharges are passed through liquid air. In this case, the solid trioxide separates as a pale blue amorphous powder which remains as a residue after the air has all evaporated. The solid melts between -103° and -111° to a deep indigo blue liquid, which commences to decompose immediately the temperature rises above -21° .

The composition of nitrogen trioxide has been established by passing the products of its decomposition over heated copper, and weighing the copper oxide and the nitrogen (p. 634). It is supposed to be represented either by—



The former agrees best with the fact that nitrogen trioxide forms nitrous acid in contact with a little cold water: $\text{N}_2\text{O}_3 + \text{H}_2\text{O} = 2\text{HNO}_2$. Hence the gas is sometimes called nitrous anhydride. If the water is warm the nitrous acid decomposes, and nitric oxide and nitric acid are formed. According to H. B. Baker (1907), if the liquid be thoroughly dried, it vaporizes without dissociation, forming a gas which has a vapour density never below 76 ($\text{H}_2 = 2$), but generally much above that number (N_2O_3). If this be confirmed, it is the only direct evidence we have of the existence of the gaseous nitrogen trioxide. The freezing-point method of determining the molecular weight in acetic acid furnished W. Ramsay (1888) with numbers between 80.9 and 92.7—theory for N_2O_3 requires 76.02. Hence the liquid is partly polymerized. Many reactions formerly said to yield the trioxide really give a mixture of nitric oxide and nitrogen peroxide.

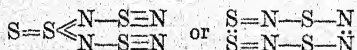
¹ The gas obtained by the action of nitric acid upon arsenic trioxide (p. 640) is a mixture of nitric oxide and nitrogen peroxide in nearly the right proportions. The same remark applies to the gas formed by the decomposition of a 20 per cent. solution of sodium nitrite with concentrated sulphuric acid.

Recapitulation.—Before dismissing the nitrogen oxides it will be found instructive to draw up a table of their comparable properties—omitting the hexoxide :

TABLE XXXV.—PROPERTIES OF THE NITROGEN OXIDES.

	Nitrogen monoxide.	Nitric oxide.	Nitrogen trioxide.	Nitrogen peroxide.	Nitrogen pentoxide.
Formula	N_2O	NO	N_2O_3	N_2O_4	N_2O_5
Atomic ratio N : O .	2 : 1	2 : 2	2 : 3	2 : 4	2 : 5
State of aggregation	Gas	Gas	Gas	Liquid	Solid
Colour of gas . . .	Colourless	Colourless	Reddish brown	Depends on temperature	White
Melting point . . .	-102.7°	-167°	-111°	-9°	c. 30°
Boiling point . . .	-89.8°	-150°	—	$+26^\circ$	c. 47°
Vapour density ($H_2 = 2$)	44	29.88	Decomposes	Varies with temperature	—
Corresponding acid	HNO	None	HNO_2	None	HNO_3

Nitrogen sulphides.—The sulphides have not been so thoroughly investigated as the oxides. When a solution of dry ammonia in benzene is treated with sulphur chloride, or when liquid ammonia acts on sulphur, long orange-red monoclinic crystals of nitrogen tetrasulphide can be obtained. The crystals have a characteristic smell, and decompose explosively at about 185° , although they may be sublimed without decomposition in vacuo at 100° . The molecular weight by freezing and boiling point methods—solution in carbon disulphide or benzene—corresponds with the formula N_4S_4 . It is decomposed by cold water, forming ammonia and sulphuric and thiosulphuric acids. It forms addition compounds with chlorine, bromine, sulphur chloride, etc., e.g. N_4S_4Cl , $N_4S_4Br_4$, $N_4S_4Br_6$. When it is sublimed in vacuo over silver gauze, a blue polymeride, N_4S_4 , is formed. The constitution of the sulphide is unknown, but it has been referred to



A deep red liquid is formed when the tetrasulphide is treated with carbon disulphide at 100° . The liquid solidifies between 10° and 11° to a mass resembling iodine. The liquid is thought to contain nitrogen pentasulphide, N_2S_5 .

Questions.

1. How would you prove that nitrous acid may act both as an oxidizing and as a reducing agent?—*London Univ.*
2. Give illustrations of the behaviour of different metals towards nitric acid.—*Aberdeen Univ.*
3. Calculate the percentages of nitrogen and oxygen in the oxides of nitrogen and show what fundamental law of chemical combination they illustrate.—*Princeton Univ., U.S.A.*
4. 30 c.c. of a compound of nitrogen and oxygen were exploded with an equal volume of hydrogen; after the explosion 30 c.c. of a gas which was pure

nitrogen remained. What is the composition (by volume) of the compound? Show from the experimental evidence that the gas could not have been a mixture of oxygen and nitrogen.—*Univ., North Wales.*

5. What is the action of concentrated nitric acid on (a) tin; (b) iodine; and (c) sulphurous acid solution?—*St. Andrews Univ.*

6. Knowing that ammonium nitrate when heated gives nitrous oxide, what gas would you expect to be produced when ammonium nitrite is heated?—*R. Galloway.*

7. What is observed and what products are obtained when each of the following nitrates is heated until any decomposition which occurs is complete: Copper nitrate, mercuric nitrate, ammonium nitrate, sodium nitrate? How could nitrogen be isolated from one of these nitrates?—*Sheffield Univ.*

8. The weight of a shilling (British coinage containing 7.5 per cent. of copper) is 6 grams. How would you prepare pure silver nitrate from it, and how much would it yield theoretically?—*Oxford Senior Locals.*

9. Outline the operations necessary for the production of liquid nitrogen peroxide from dry lead nitrate. Describe carefully the changes in colour observed when nitrogen peroxide is heated until it becomes colourless, and subsequently cooled. What explanation can be given of these changes?—*Sheffield Univ.*

10. Give the preparation and properties of nitric oxide. Show how the formula of the gas may be deduced from the following data:—15.6 c.c. of the gas passed over heated copper give 7.8 c.c. of nitrogen; the weight of nitric oxides which fills a certain globe is 3.75 grams, the weight of an equal volume of hydrogen being 0.25 gram.—*Cambridge Senior Locals.*

11. Describe two examples of reactions which may be brought about by (a) the electric current; (b) the silent electric discharge; (c) the electric spark; and in any two instances describe exactly how the reaction may be carried out experimentally.—*London Univ.*

12. Describe an analysis of nitrous oxide made by exploding the gas with an excess of hydrogen, and the residual hydrogen with oxygen; and show how the molecular formula of nitrous oxide is inferred from the volumes of gas at the successive stages.—*New Zealand Univ.*

13. Upon what changes does the oxidizing action of nitric acid depend? Mention three examples of its action.—*London Univ.*

14. We see the formula of nitrogen peroxide sometimes written NO_2 and sometimes as N_2O_4 . Which is the correct, and on what experimental evidence is your opinion based?—*Punjab Univ.*

15. Chlorine and nitric acid are called *oxidizing agents*, and sulphuretted hydrogen and hydriodic acid are called *reducing agents*. Explain the terms in italics and give examples illustrating the oxidizing and reducing action of these bodies.—*Cambridge Senior Locals.*

16. Plot the specific gravity of solutions of nitric acid with their concentrations from the following data:

HNO_3 (per cent.)	3.7	10.68	30.88	59.39	72.17	82.90	99.67
Specific gravity	1.020	1.060	1.190	1.370	1.430	1.470	1.520

Indicate the percentage composition of solutions of specific gravity 1.170 and 1.500.

Answers: 27.88 and 94.09.

CHAPTER XXVIII

COMPOUNDS OF NITROGEN AND HYDROGEN

§ 1. Ammonia—Occurrence and Preparation.

Molecular weight, $\text{NH}_3 = 17.03$. Melting point, -78° ; boiling point, -33.5° ; critical temperature, $132.9^\circ \pm 0.1^\circ$. Vapour density ($\text{H}_2 = 2$), 16.97; (air = 1), 0.5971. One litre weighs 0.77079 gram under normal conditions

History.—Ammonia was known to the early chemists, and Geber describes the preparation of ammonium chloride by heating urine and common salt. Hence the alchemists' term—*spiritus salis urinae*. Ammonium chloride was first brought to Europe from Egypt, where it was prepared from the "soot" obtained by burning camel's dung. The name ammonia seems to be connected somehow with the Egyptian sun-god—Ra Ammon; ammonium salts must have been known to the early Egyptian priests. The term *sal ammoniac* was one of the early names for ammonium chloride; the equivalent term *sal armoniacum* which appears in the translations of Geber's writings, and which was used for some time afterwards, was probably a mis-spelling, since the term "salt of Armenia"—*sal armoniacum*—was applied to common salt and to native sodium carbonate. In 1716 J. Kunckel mentioned the formation of ammonia during fermentation. S. Hales (1727) noticed that when lime was heated with sal ammoniac in a retort arranged to collect the gas over water, no gas appeared to be given off; on the contrary, water was sucked into the retort; when J. Priestley (1774) tried the experiment with a mercury gas trough, he obtained ammonia gas which he called "alkaline air." C. L. Berthollet (1785), H. Davy (1800), and others established the composition of the gas.

Occurrence.—Small quantities of ammonia occur in atmospheric air and in natural waters. It is produced by the action of putrefying bacteria (p. 614) on organic matter in the soil, etc. The odour of ammonia can often be detected near stables. Ammonium salts are also deposited on the sides of craters and fissures of the lava streams of active volcanoes; and with boric acid in the fumaroles of Tuscany.

Preparation.—Ammonia can be obtained by reducing nitric acid, nitrates, or nitrites with nascent hydrogen (p. 625). For instance, by boiling a mixture of 25 grams of sodium hydroxide in 70 c.c. of water, 20 grams of zinc or aluminium turnings, a piece of bright sheet iron, and 5 grams of nitric acid in a half-litre flask, ammonia gas is slowly evolved: $4\text{Zn} + 8\text{NaOH} + \text{HNO}_3 = 4\text{Zn}(\text{ONa})_2 + 3\text{H}_2\text{O} + \text{NH}_3$. The reduction of the nitric acid by the hydrogen liberated by the action of zinc on the alkaline solution (p. 113) is rather too slow to make this a suitable process

for making ammonia for "lecture table" experiments with the gas. The reaction, however, is rather important because upon it is based a method for determining the quantity of nitrites and nitrates in potable water. The formation of ammonia during the rusting of iron has been confirmed by numerous observers since C. Bourdelin (1683) first mentioned the fact. The ammonia seems to be formed either by the action of nascent hydrogen on the nitrogen of the air: $2\text{Fe} + 6\text{H}_2\text{O} + \text{N}_2 = 2\text{Fe}(\text{OH})_3 + 2\text{NH}_3$, or else by the reduction of nitrates commonly present in air: $8\text{Fe} + 15\text{H}_2\text{O} + 3\text{HNO}_3 = 8\text{Fe}(\text{OH})_3 + 3\text{NH}_3$. The reduction of nitric acid by zinc is an illustration of the last-named process. Aluminium amalgam reduces concentrated solutions of the nitrites or nitrates to ammonia: $3\text{KNO}_3 + 8\text{Al} + 21\text{KOH} = 3\text{NH}_3 + 8\text{K}_3\text{AlO}_3 + 6\text{H}_2\text{O}$.

When ammonia gas is required for the laboratory, it may be obtained from a cylinder of liquid ammonia; by boiling commercial aqua ammonia in a flask—Fig. 214, or by heating an intimate mixture of commercial ammonium chloride or ammonium sulphate with twice its weight of quicklime, CaO , or slaked lime, $\text{Ca}(\text{OH})_2$. The reaction is represented: $2\text{NH}_4\text{Cl} + \text{Ca}(\text{OH})_2 = \text{CaCl}_2 + 2\text{H}_2\text{O} + 2\text{NH}_3$. Ammonia combines with the ordinary drying agents—calcium chloride, sulphuric acid, phosphorus pentoxide—and accordingly these agents must not be used with the idea of drying the gas. A tower of quicklime, indicated in Fig. 214, is generally employed. If a mixture of ammonium chloride and quicklime is to be heated, a copper flask, without the tube funnel, Fig. 214, is preferable to glass since steam is liable to condense on the walls of the flask, trickle down, and crack the hot glass. Otherwise the disposition of the apparatus is similar.

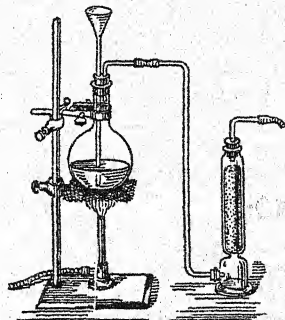


FIG. 214.—Preparation of Ammonia.

Ammonia is formed during the action of water on some of the metallic nitrides—compounds of the metals, magnesium, calcium, lithium, beryllium, aluminium, etc., with nitrogen. For instance, with hot water: $\text{Mg}_3\text{N}_2 + 6\text{H}_2\text{O} = 3\text{Mg}(\text{OH})_2 + 2\text{NH}_3$; and $\text{AlN} + 3\text{H}_2\text{O} = \text{Al}(\text{OH})_3 + \text{NH}_3$, etc. Lithium gives the most satisfactory results, but technical difficulties prevent the industrial application of lithium. Aluminium nitride is used in Serpek's process for ammonia. The action of superheated steam on calcium cyanamide (*q.v.*) also furnishes ammonia: $\text{CaCN}_2 + 3\text{H}_2\text{O} = \text{CaCO}_3 + 2\text{NH}_3$. The yield is 99 per cent. of that indicated by this equation. The process appears likely to have a commercial future.

Ammonia is formed during the putrefaction of organic compounds, and when nitrogenous compounds—coal, leather, bones, etc.—are heated in closed vessels. The old term for ammonia—*spiritus cornus cervi* (*spirits of hartshorn*)—refers to an old custom of preparing ammonia by heating hoofs and horns of stags, etc., in closed vessels. The formation of ammonia by heating nitrogenous compounds in closed vessels is particularly noticeable if the organic matter be heated with soda lime—that is, quicklime slaked with a concentrated solution of sodium hydroxide. In fact, this is

the principle of one of the standard methods for the determination of nitrogen in organic matter. The ammonia is absorbed in sulphuric acid. Most of the ammonia of commerce is derived from the ammoniacal liquid obtained as a by-product in the manufacture of coal gas (*q.v.*), in the manufacture of iron in the blast furnace (*q.v.*), and of producer gas. This liquid is boiled with milk of lime, and the ammonia which is evolved is absorbed by dilute sulphuric acid— $2\text{NH}_3 + \text{H}_2\text{SO}_4 = (\text{NH}_4)_2\text{SO}_4$. The ammonium sulphate so obtained crystallizes without evaporation when a sufficient concentration has been attained. It is removed by perforated ladles and recrystallized, or heated with milk of lime, and the resulting gas absorbed in distilled water to form the *aqua ammonia* of commerce. The solution of ammonia gas in water is conventionally styled "ammonia."

Nitrogen and hydrogen unite directly when a mixture of the two gases is subjected to electric sparks (Fig. 169). The reaction appears to stop when about 2 per cent. of ammonia has been formed, and 98 per cent. of gas remains uncombined. The converse result is obtained if ammonia gas be exposed to the electric sparks, 98 per cent. decomposes. The reaction is therefore reversible: $2\text{NH}_3 \rightleftharpoons \text{N}_2 + 3\text{H}_2$. If water or acid be present, the ammonia is absorbed as fast as it is formed, and the reaction proceeds to an end. All the nitrogen and hydrogen combine. The two gases nitrogen and hydrogen also combine if heated in the presence of finely divided iron, manganese, uranium, tungsten, zinc, osmium, etc. The reaction is exothermal: $\text{N} + 3\text{H} = \text{NH}_3 + 11.4 \text{ Cals.}$ The amount of ammonia gas in equilibrium decreases as the temperature rises. This is just the opposite effect to that observed in the oxidation of nitrogen to nitric oxide. At 1000° the ammonia is almost completely decomposed. Thus, at atmospheric pressure, and at 30 atmospheres pressure :

Temperature . . .	700°	801°	901°	974°
Ammonia (1 atm.) .	0.022	0.012	0.007	0.005 per cent. per volume.
Ammonia (30 atms.)	0.654	0.344	0.207	0.15 per cent. per volume.

The amount of ammonia formed is considerably increased if the pressure be raised. G. Claude works at a higher pressure—1000 atmospheres. If the gases be kept in circulation so that the ammonia can be removed from the zone of the reaction, F. Haber (1906), working at 185 atmospheres pressure, obtained 90 grams of liquid ammonia per hour from a small model apparatus. Good results have also been obtained with uranium carbide at 550° as a catalytic agent, and at 750° , and a pressure of 175 atmospheres. The process has recently been taken up by the Badische Anilin und Sodafabrik for exploiting commercially, and the use of iron or molybdenum as a catalyst at 600° has been patented.

Uses.—Ammonia is used in refrigerating machines; as a cleansing agent on account of its property of dissolving greases; in the manufacture of soda by the Solvay process; in chemical operations where a volatile alkali is needed; etc. A carboy of ammonia (sp. gr. 0.88) holds 88–90 lbs. The commercial ammonia sells at about 5*d.* per lb., and the pure at 6½*d.* per lb. The commercial aqueous ammonia can be purchased with from 20 to 35 per cent. of NH_3 by weight. That with a specific gravity 0.88 contains about 35 per cent. of NH_3 . The chief impurities are pyridine, chlorides, sulphates, carbonates, phosphates, lime and magnesia, and sulphides.

§ 2. The Properties of Ammonia.

Ammonia is a colourless gas with a pungent odour. If inhaled suddenly, it will bring tears to the eyes, if large quantities be inhaled, suffocation may ensue. Ammonia is a little more than half as heavy as air, and consequently, the gas is collected, like hydrogen, by the downward displacement of air.

Action of water.—The gas is extremely soluble in water: one volume of water at 0° and 760 mm. dissolves 1298 volumes of gas, and at 20° , 710 volumes. The gas can all be removed from its aqueous solution by boiling. The great solubility of ammonia in water is illustrated by means of the apparatus indicated in Fig. 213, but an ordinary flask will do. If a jar of dry ammonia be collected over mercury, and carried on a dish of mercury into a tray of water, and the dish of mercury removed below the surface of water, the absorption of ammonia is so rapid that the cylinder is often broken.¹ The specific gravity of the solution of ammonia in water gradually decreases as the concentration of the ammonia increases. The freezing curves of mixtures of ammonia and water by F. F. Rupert (1909), Fig. 215, show three eutectics and the existence of two hydrates: ammonia semihydrate, $2\text{NH}_3 \cdot \text{H}_2\text{O}$, which forms small needle-like crystals melting at -79° ; and ammonia monohydrate, $\text{NH}_3 \cdot \text{H}_2\text{O}$, corresponding with the so-called ammonium hydroxide, NH_4OH . The latter compound forms needle-like transparent crystals melting at -79° . Solutions of ammonia containing 25 to 60 per cent. of ammonia are very viscid at temperatures below -60° .

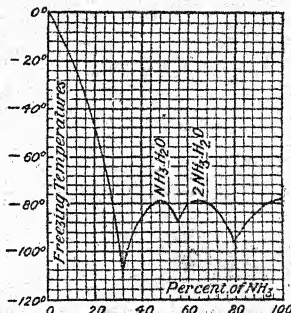


Fig. 215.—Freezing Curves of Aqueous Ammonia.

Considerable heat is evolved during the solution of the gas: $\text{NH}_3 + \text{Aq} = \text{NH}_3\text{Aq} + 8.4 \text{ Cals.}$ If a rapid current of air be driven through a cold solution of ammonia in water, the heat absorbed as the ammonia is expelled from the solution will reduce the temperature so as to freeze a small globule of mercury. The production of cold is best demonstrated by blowing air through a solution of ammonia standing on a few drops of water on a block of wood. The beaker will soon be frozen to the block of wood. Carre's ice machine is a time-worn application of the principle.

Refrigeration.—The heat of evaporation of liquid ammonia is 5.7 Cals. at -33° . This means that 17 grams of liquid ammonia at -33° requires 5.7 Cals. of heat before it can pass into a gas at -33° . Otherwise expressed, when the gas is liquefied, heat is liberated; and conversely, heat is absorbed when the liquid is vaporized. If, therefore, liquid ammonia be evaporated, a relatively large amount of heat is absorbed from its surroundings. Advantage is taken of this fact in the preparation of artificial ice, cold storage, etc. Ammonia gas is liquefied by

¹ Consequently the hand is protected by holding the cylinder with a piece of cloth.

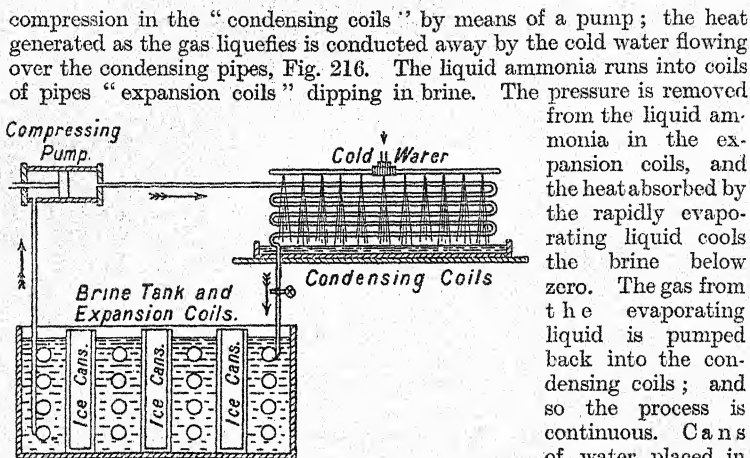


FIG. 216.—Refrigeration Plant (Diagrammatic).

storage rooms, the cold brine circulates in coils near the ceiling of the room to be cooled, and returns to the cooling tank. Sulphur dioxide and carbon dioxide are also employed for refrigeration; but ammonia is more generally used.

Double Compounds.—Aqueous solutions of ammonia often act very like solutions of the alkali hydroxides. Thus, with ferric chloride: $\text{FeCl}_3 + 3\text{NH}_4\text{OH} = \text{Fe}(\text{OH})_3 + 3\text{NH}_4\text{Cl}$; in other cases, the ammonia reacts with the salt to form soluble or insoluble complex or double salts. Ammonia is absorbed by calcium chloride, zinc chloride, silver chloride, etc., forming double or complex compounds. For instance, with silver chloride below 15° , ammonia forms $\text{AgCl} \cdot 3\text{NH}_3$; above 20° , $2\text{AgCl} \cdot 3\text{NH}_3$; with calcium chloride, $\text{CaCl}_2 \cdot 2\text{NH}_3$; $\text{CaCl}_2 \cdot 4\text{NH}_3$; $\text{CaCl}_2 \cdot 8\text{NH}_3$; and with aluminium chloride, compounds with 1, 3, 5, 6, and 9 molecules of ammonia have been reported. Hundreds of complex compounds of the mercury salts with ammonia have been reported.

Liquid and solid ammonia.—It is sometimes claimed that the first known liquefaction of a gas was the condensation of ammonia by M. van Marum about 1792. There was, however, L. Clouet and G. Monge's liquefaction of sulphur dioxide before 1782. If the compound of silver chloride with ammonia be heated in one leg of a V-shaped hermetically closed tube immersed in water, Fig. 109, and the other leg immersed in a freezing mixture—say calcium chloride and ice—the ammonia gas condenses in the cold leg of the V-tube to a colourless limpid liquid. Like water, liquid ammonia is a bad conductor of electricity. The liquid boils at -33.5° , and solidifies to white transparent crystals at -78° . The specific heat of liquid ammonia and the heat of fusion of the solid are greater than the corresponding constants for water; and the heat of volatilization and the critical temperature and pressure are higher than for any other liquid except water.

Oxidation of ammonia.—Ammonia is a non-supporter of ordinary

combustion and it is incombustible in air. It burns in oxygen, forming nitrogen, water, and small quantities of ammonium nitrate, and nitrogen peroxide. If ammonia be mixed with oxygen, say, by bubbling a stream of oxygen through a small quantity of concentrated ammonia warmed in a flask, the gas issuing from the flask can be ignited; it burns with a yellow flame. In a few moments the solution in the flask will be too dilute to show the flame. By sending a jet of ammonia into the air-holes of a Bunsen's burner, the flame will be found to expand, and acquire a yellow tinge. The effect is shown better by delivering a jet of ammonia into the centre of the tube of a Bunsen's burner. If a stream of oxygen be sent into a cylinder fitted as shown in Fig. 217, and a stream of ammonia be sent into the same cylinder through a wide glass tube, the ammonia can be ignited, and it will burn with a yellowish flame. In K. Kraut's experiment (1865), a stream of oxygen is sent through a concentrated solution

of ammonia in a beaker in which is suspended a spiral of thin ($\frac{1}{4}$ mm.) platinum wire (recently ignited), Fig. 218. If the current of oxygen be very slow, the platinum wire will glow red hot, and the beaker will soon be filled with brown fumes of nitrogen peroxide. If the current of oxygen be faster, a small explosion will occur every now and again: the first explosion will be stronger than the second, the second stronger than the third, etc., and the solution in the beaker will be found to contain both ammonium nitrite and nitrate. Here the platinum acts as a catalytic agent (p. 160). If a mixture of ammonia and air be passed through a tube (Fig. 210) containing hot platinized asbestos, nitric oxide is formed: $4\text{NH}_3 + 5\text{O}_2 = 6\text{H}_2\text{O} + 4\text{NO}$. The air is driven through a warm solution of ammonia and then passed over the platinized asbestos. The nitric oxide is subsequently oxidized to nitrogen peroxide, and then to nitric acid. The experiment is due to Kuhlmann. If the current of gases be slow, the temperature of the platinum is raised so high that the nitrogen oxides are decomposed. According to W. Ostwald, the best temperature for a maximum yield is about 300° . This principle is applied at a works near Bochum for the manufacture of nitric acid from ammonia. A mixture of ammonia and oxygen explodes violently when ignited: $4\text{NH}_3 + 3\text{O}_2 = 2\text{N}_2 + 6\text{H}_2\text{O}$. An aqueous solution of ammonia is oxidized to ammonium nitrite and nitrate and hydrogen peroxide by ozone. An ammonium peroxide, $(\text{NH}_4)_2\text{O}_2$, analogous with sodium peroxide, has been prepared. It dissociates at temperatures exceeding -10° .

Action of ammonia on the metals.—When magnesium is heated in

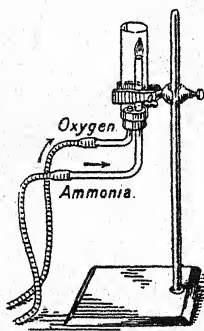


FIG. 217.—Combustion of Ammonia.

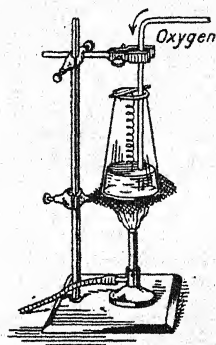
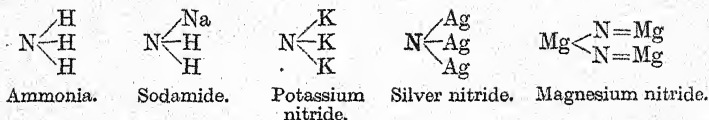


FIG. 218.—Kraut's Experiment.

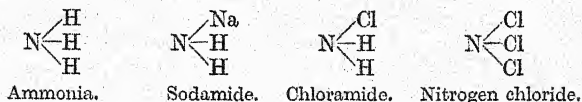
ammonia gas, it forms magnesium nitride, Mg_3N_2 . As indicated on p. 649, the nitrides give ammonia when treated with water. Several other nitrides are known, e.g. Ca_3N_2 , AlN , etc. When ammonia gas is passed over hot sodium, a compound NH_2Na —sodamide—is formed. Sodamide is made by heating, say, 3 grams of metallic sodium in a deep nickel boat, in a combustion tube, or iron vessel, between 300° and 400° in a current of ammonia dried by passing it through a tower of soda lime. Hydrogen mixed with the excess of ammonia escapes. The reaction is represented: $2NH_3 + 2Na = 2NH_2Na + H_2$. The amides of potassium, sodium, etc., are usually decomposed by water, forming ammonia, and the hydroxides of the metals: $RbNH_2 + H_2O = RbOH + NH_3$. The existence of compounds like



is sometimes cited as evidence that ammonia is a feeble tribasic acid. This statement turns on what definition of an acid be adopted. In any case it has been shown, p. 167, that a compound may be an acid under one set of conditions, and a base under another set.

Action of chlorine on ammonia.—Chlorine decomposes ammonia gas, forming nitrogen (*q.v.*) and hydrogen chloride. The latter unites with the remaining ammonia, forming ammonium chloride. The complete reaction with *an excess of ammonia* is therefore represented: $8NH_3 + 3Cl_2 = 6NH_4Cl + N_2$. If ammonium chloride be treated with an *excess of chlorine*—say, by inverting a jar of chlorine over a warm (30° – 40°) concentrated solution of ammonium chloride—yellow oily drops of nitrogen chloride, NCl_3 , are formed: $NH_4Cl + 3Cl_2 = NCl_3 + 4HCl$. The reaction is endothermic: $N + 3Cl = NCl_3 - 42$ Cals. Hence, in preparing nitrogen by the action of chlorine on aqueous ammonia, if the current of chlorine be continued after all the ammonia has been converted into ammonium chloride, the unmanageable explosive, nitrogen chloride, may be produced. When an ammonium chloride solution is electrolyzed, the ammonium cation breaks down into ammonia and hydrogen at the cathode, and the chloride liberated at the anode reacts with the ammonium chloride forming drops of the same violent explosive, nitrogen chloride. Nitrogen chloride is one of the most explosive substances known. It was discovered by P. L. Dulong in 1811, who continued to work on it after it had caused him the loss of three fingers and one eye. It explodes violently when in contact with many organic substances; exposure to sunlight; and also spontaneously. Nitrogen chloride is also formed when a solution of ammonium chloride is electrolyzed; it is also formed when a lump of ammonium chloride is suspended in a concentrated solution of hypochlorous acid, $HOCl$, and since nitrogen chloride is hydrolyzed by water into ammonia and hypochlorous acid, the reaction $NH_3 + 3HOCl \rightleftharpoons 3HOH + NCl_3$, must be reversible. The compound is more stable when it is dissolved in certain solvents, say carbon tetrachloride, CCl_4 . Concentrated hydrochloric acid and ammonia both decompose nitrogen chloride, forming ammonium chloride and chlorine in the one case, and ammonium chloride

and nitrogen in the other. Monochloramide— NH_2Cl —is obtained by adding, say, 50 c.c. of a solution of 3.7 grams of sodium hypochlorite—free from an excess of chlorine—to 100 c.c. of an aqueous solution containing 0.85 gram of ammonia. The liquid ceases to smell of ammonia, and in its place, a penetrating smell of monochloramide, NH_2Cl , is developed: $\text{NH}_3 + \text{NaOCl} = \text{NaOH} + \text{NH}_2\text{Cl}$. Some nitrogen is at the same time evolved owing to the decomposition of the monochloramide: $3\text{NH}_2\text{Cl} = \text{N}_2 + \text{NH}_4\text{Cl} + 2\text{HCl}$. The relation of these compounds to ammonia will appear from the graphic formulæ:



Nitrogen iodide, $\text{N}_2\text{H}_3\text{I}_3$.—The chocolate coloured amorphous powder of nitrogen iodide obtained by the action of aqueous solution of ammonia upon iodine—either solid or in solution—was once thought to be NI_3 , analogous with nitrogen chloride, NCl_3 . The compound decomposes readily both in diffused daylight, and in the presence of the liquids employed in washing the powder prior to analysis. An exact determination of the composition of the pure substance is therefore difficult, and different results have been obtained with slight variations in the mode of preparation. Many have demonstrated that the compound contains hydrogen; and, in consequence, the formula has been variously given as $\text{N}_2\text{H}_3\text{I}_3$, NH_2I , NHI_2 , NH_3I_2 , etc.

Nitrogen iodide is prepared in a state of purity by the action of ammonia upon an aqueous solution of potassium hypoiodite, KIO , in the presence of excess of potassium hydroxide. If the solutions employed be not too concentrated, nitrogen iodide separates in brownish-red crystals with the ultimate composition: $\text{N}_2\text{H}_3\text{I}_3$. This may mean that the substance is a compound of nitrogen iodide and ammonia— $\text{H}_2\text{N} : \text{NI}_3$. The reactions just indicated throw some light on the constitution. In the first case, ammonium hypoiodite is probably first formed by the action of ammonia on the solution of iodine: $\text{I}_2 + 2\text{NH}_4\text{OH} = \text{NH}_4\text{IO} + \text{NH}_4\text{I} + \text{H}_2\text{O}$; and with potassium hypoiodite: $\text{KIO} + \text{NH}_4\text{OH} = \text{NH}_4\text{IO} + \text{KOH}$. The unstable hypoiodite spontaneously decomposes, producing the compound in question: $3\text{NH}_4\text{IO} \rightleftharpoons \text{N}_2\text{H}_3\text{I}_3 + \text{NH}_4\text{OH} + 2\text{H}_2\text{O}$. This reaction is to some extent reversible, and a small quantity of ammonium hypoiodite is reproduced when nitrogen iodide is suspended in dilute ammonia and exposed to light.

The main action during the decomposition of nitrogen iodide in light is a simple decomposition into nitrogen gas and hydrogen iodide: $\text{N}_2\text{H}_3\text{I}_3 = \text{N}_2 + 3\text{HI}$. The action is further complicated by the interaction of the hydrogen iodide with the still undecomposed solid. Dry nitrogen iodide exposed to light behaves in a similar manner. Nitrogen iodide is rapidly decomposed by acids and alkalis. Although moist nitrogen iodide can be handled without much danger of explosion, the dry compound is very explosive. Explosions are said to have been produced as a result of the shock of a falling dust-particle, and by a fly walking over the dry powder.

Ammonia as a base.—One of the most striking properties of ammonia, chemically speaking, is the basic character of its aqueous solution. The aqueous solution turns red litmus blue, yellow tumeric paper brown, conducts electricity, and in general reacts like a base. Indeed it is supposed that a molecule of water combines with a molecule of ammonia to form a solution of ammonium hydroxide: $\text{NH}_3 + \text{H}_2\text{O} = \text{NH}_4\text{OH}$. Hence aqueous ammonia is sometimes called ammonium hydroxide. If the aqueous solution of ammonia be neutralized with an acid—nitric, sulphuric or hydrochloric acid—the corresponding ammonium salt is formed—ammonium nitrate, NH_4NO_3 ; ammonium sulphate, NH_4HSO_4 , or $(\text{NH}_4)_2\text{SO}_4$; ammonium chloride, NH_4Cl . It will be observed that we are here dealing with a univalent radicle, NH_4 , which is called ammonium. The assigning of the name does not make this radicle any less hypothetical than if it were nameless. The fact that ammonium appears to form a series of salts closely analogous with the salts of sodium and potassium has instigated many to seek for a compound, NH_4 , with a corporeal existence.

§ 3. Ammonium Amalgam.

When a little ^{sodium}mercury amalgam, *A*, Fig. 219 (containing about one per cent. of sodium or potassium) is placed in a solution of ammonium chloride, the mercury swells up into a frothy mass, illustrated in *B*, Fig. 219, thirty times its original volume. The inflated mass can be compressed or expanded by raising or lowering the pressure. Again, if mercury be brought into a concentrated aqueous solution of ammonium chloride, and a current of electricity be passed through the solution in such a way that the cathode dips into mercury (Fig. 220), the mercury

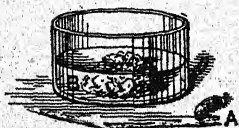


FIG. 219.—Ammonium Amalgam.

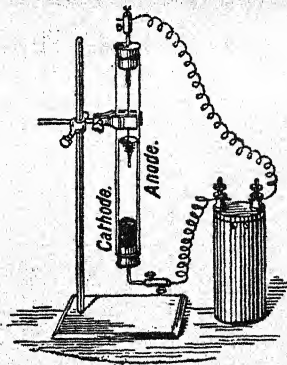


FIG. 220.—Ammonium Amalgam.

swells up in a similar manner. If the temperature be kept below 0° , the amalgam shows little tendency to inflation.

H. Moissan assumed that the amalgam is a compound of ammonia with mercury hydride since sodium amalgam containing sodium hydride readily reacts with aqueous ammonia with an increase in volume, while sodium amalgam alone gives hydrogen without appreciably changing its volume. Sodium amalgam, like that of ammonium, swells up in contact with a solution of ammonium salts. The argument is inconclusive. Others consider that the amalgam is a mere solution of ammonia and hydrogen in

mercury. Against this view it is urged that (1) neither of these gases (ammonia or hydrogen) alone or mixed together will dissolve in this manner; (2) the gases bear to one another the ratio $2\text{NH}_3 : \text{H}_2$. There is not much in this argument, because if electrolytic gas were bubbled through a soap-solution, the gas obtained from the froth would have a constant composition. (3) When ammonium amalgam prepared at 0° is brought into contact with cold solutions of copper, cadmium, or zinc sulphate, some of the metal is precipitated and ammonium sulphate is formed: it is supposed that: $2\text{NH}_4 + \text{CuSO}_4 = \text{Cu} + (\text{NH}_4)_2\text{SO}_4$. At higher temperatures no precipitation occurs. Neither zinc nor cadmium are reduced to the metal by hydrogen or ammonia. (4) At low temperatures, 0° , the NH_4 of ammonium amalgam can expel sodium from its amalgams; similarly the alkali and alkaline earth metals are displaceable from their amalgams. Hence ammonium amalgam here behaves like other metal-amalgams. The conclusion is that the so-called ammonium amalgam is a true amalgam and that the constituent NH_4 has the properties of an alkali metal, for its behaviour is quite analogous with that of potassium in potassium amalgam, or sodium in sodium amalgam. The solution of ammonium NH_4 in mercury decomposes into mercury, hydrogen, and ammonia when warmed above 0° . These gases, entangled with the mercury, cause the frothing.

The sulphonium bases, $\text{SR}_3'\text{OH}$, of organic chemistry are related to hydrogen sulphide, H_2S , similar to the way ammonium hydroxide, NH_4OH , is related to ammonia, NH_3 .

§ 4. Ammonium Salts.

Elements and compounds totally unlike in almost every particular display a great degree of similarity by suitably modifying their attractive power by combining them with suitable radicles.—G. MARTIN.

No two elements are more unlike chemically than one of the alkali metals and nitrogen, but by combining nitrogen with four atoms of hydrogen, the properties of nitrogen are completely transformed, and the resulting radicle, ammonium, has properties which bear a most perfect analogy in a whole series of compounds and reactions with those of the alkali metals. The ammonium compounds are therefore classed with the alkali metals. The ammonium salts are usually very soluble in water, and when the solutions are boiled, partial decomposition occurs. The solution, originally neutral, may become acid owing to the volatilization of more ammonia than acid. When the ammonium salts are heated with the fixed alkalies—potassium or sodium—or calcium hydroxides, or calcium oxide, the ammonia is volatilized. Hence the old term *volatile alkali* for ammonia.

Identification of ammonia and ammonium salts.—(1) When the salts are heated with an alkali or lime, the characteristic smell of ammonia is obtained. (2) With hydrochloroplatinic acid, they give a yellow precipitate of ammonium chloroplatinate. (3) Nessler's reagent gives a yellow coloration. The intensity of the tint with a given concentration is nearly proportional to the amount of ammonia present.

Ammonium sulphate.—This salt is usually made from "gas liquor" as indicated on p. 873. It is also formed by the neutralization of ammonia with dilute sulphuric acid. It is used principally as a fertilizer, and also in the manufacture of ammonium compounds. Hence, in the words of

A. W. Hofmann, after the lapse of countless ages, the nitrogen of petrified fern forests (*i.e.* coal) is resuscitated in the ammoniacal liquor of the gas works to vegetate once more, and increase the product from our cornfields.

Ammonium nitrate.—Four different types of crystals are known, each type has a definite transition temperature. The ordinary crystals are rhombic, and isomorphous with potassium nitrate. The heat of solution is: $2\text{NH}_4\text{NO}_3 + 400 \text{ Aq} = -12.6 \text{ Cals.}$ If 60 parts of the salt be dissolved in 100 parts of water at 13° , the temperature of the liquid falls to about -13° ; and if the water be at 0° , the temperature of the liquid falls to about -16° . Hence a mixture of ice and ammonium nitrate is a valuable freezing mixture for reducing the temperature below the freezing point of water. A mixture of ice and common salt is often used for the same purpose. When heated, the salt melts at about 166° , and the solution appears to boil at 180° , this is because the salt is decomposing, forming nitrous oxide. If sharply heated, by throwing a crystal on a red-hot surface, the salt deflagrates, and it may be exploded by means of a detonator. The explosion is more violent if the nitrate be mixed with some oxidizable substance, say naphthalene. This salt is used chiefly in the preparation of nitrous oxide; and in the manufacture of fireworks and safety explosives—so called because the temperature of the explosion is low compared with that of gunpowder. For instance, the explosive "ammonite" is said to contain between 80 and 90 per cent. of this salt.

Ammonium nitrite.—This salt is made by saturating an aqueous solution of ammonia with nitrous acid, or by adding silver nitrite to a solution of ammonium chloride. When heated the solution decomposes into nitrogen and water, so that the crystals cannot be prepared by evaporation in the ordinary manner. The solid can be obtained by evaporating a clear aqueous solution of the salt over sulphuric acid *in vacuo* at ordinary temperatures; or better, by adding ether to an alcoholic solution of the salt when crystals of ammonium nitrite separate. Very little gas is evolved if the salt be heated *in vacuo* below 40° ; on cooling most of the salt crystallizes; at 70° , the salt slowly decomposes and a large part sublimates.

Ammonium chloride.—If ammonia gas be brought in contact with hydrogen chloride, dense white fumes of ammonium chloride are formed:

$\text{NH}_3 + \text{HCl} = \text{NH}_4\text{Cl}.$ This can be illustrated by placing a bottle of aqueous ammonia alongside a bottle of hydrochloric acid, and blowing across the mouth of the one bottle to the other. Ammonium chloride is made by mixing an aqueous solution of ammonia and hydrochloric acid, and by passing ammonia gas into dilute hydrochloric acid. The last-named process is usually employed on a manufacturing scale. The crude product is purified by heating the solid in a large iron or earthenware pot with a dome-shaped cover, Fig. 221.

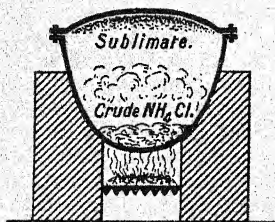


FIG. 221.—Sublimation of Ammonium Chloride.

The ammonium chloride volatilizes and the solid condenses as a white crystalline fibrous mass inside the cover. Most of the impurities remain in the vessel. The process of vaporizing a solid and condensing the

vapour back to the solid condition is called *sublimation*. Ammonium chloride is a white granular, fibrous, crystalline solid, with a sharp saline taste. It dissolves in water and at the same time lowers the temperature. The heat of solution is -3.7 Cals. The solid undergoes a change at 159° whereby it contracts in volume; ammonium bromide undergoes a similar change at 109° ; and ammonium iodide does not change. Ammonium chloride is used for charging Leclanché cells; as a constituent of soldering fluids, to protect metals from oxidation during the soldering. It is also used in galvanizing iron, and in the textile industries.

§ 5. The Dissociation of Ammonium Chloride.

The vapour density ($H_2 = 2$) of ammonium chloride at 350° is 29.04 ; and at 1040° , 28.75 . The theoretical value for the molecule NH_4Cl is 53.5 . Hence we cannot be dealing with that molecule at the temperatures named. Suppose the vapour of ammonium chloride be dissociated so that a mixture

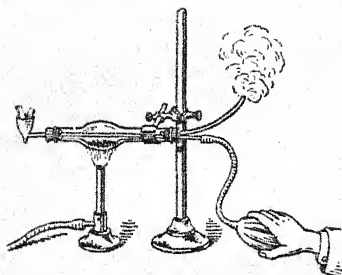


FIG. 222.

Dissociation of Ammonium Chloride.

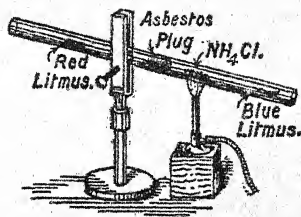


FIG. 223.

of equal volumes of ammonia and hydrogen chloride is formed corresponding with $NH_4Cl = NH_3 + HCl$. The vapour density for complete dissociation would then be 26.75 , that is, half the value for NH_4Cl —that is, $\frac{1}{2}$ of $(17 + 36.5)$. The experimental result thus shows that dissociation is nearly complete. Applying the method of p. 640, it follows that the vapour contains about 17 per cent. of ammonium chloride, and 83 per cent. of a mixture of equal volumes of ammonia and hydrogen chloride.

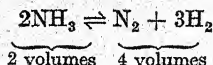
According to H. B. Baker (1894), if the vapour density be determined in a vessel of hard glass with a thoroughly dried sample of ammonium chloride, the number is quite normal, namely, 53.4 . This shows that the dry salt does not dissociate quickly enough to affect the determination. Obviously, too, the moisture is a catalytic agent which accelerates not only the dissociation, but also the formation of ammonium chloride from ammonia and hydrogen chloride. The constituent gases were dried so thoroughly by H. Tramm that they did not unite at ordinary temperatures. Confirmatory evidence that ammonium chloride dissociates when heated is as follows: (1) When the component gases are brought together at about 400° , no heat is evolved, and there is no concomitant sign of chemical combination. (2) The heat absorbed in volatilizing ammonium chloride and the heat evolved when this salt is formed from ammonia and

hydrogen chloride are the same. (3) The mixture of the component gases in the vapour can be separated by mechanical means.

L. von Bebal illustrated the dissociation of ammonium chloride under ordinary conditions by taking advantage of the difference in the speeds of diffusion of the two gases ammonia and hydrogen chloride. The process of atmolysis (p. 129) can be used. Let the stem of a clay pipe be fixed by corks in a hard glass bulb-tube with some solid ammonium chloride in the bulb, as illustrated in Fig. 222. The mouth-piece of the clay pipe is connected with a rubber blower. One cork is fitted with an exit tube. A strip of red and a strip of blue litmus paper are placed in the bowl. When the ammonium chloride has been heated for a short time, blow a very gentle current of air through the stem of the pipe. The red litmus will be coloured blue by the ammonia gas which diffuses through the porous earthenware much faster than the hydrogen chloride. The issuing vapour of ammonium chloride reddens blue litmus. Another instructive way of showing the same phenomenon is as follows: Place a little ammonium chloride near the middle of a piece of hard tube (Fig. 223), and a little lower down the tube place a piece of blue litmus paper. Place a loose plug of asbestos a little above the salt, and then a piece of red litmus paper. Heat the ammonium chloride. The ammonia being the lighter gas diffuses more quickly than the hydrogen chloride. Consequently, when the ammonium chloride is heated, the blue litmus will be reddened by the excess of slow diffusing hydrogen chloride in the lower part of the tube; and the red litmus will be blued by the ammonia which passes to the upper part of the tube before the hydrogen chloride.

§ 6. The Composition of Ammonia.

1. By explosion with oxygen.—Repeating old experiments by M. von Marum and by C. L. Berthollet (1785), if ammonia gas be sparked in an apparatus, say, Fig. 170, the volume of the gas will be nearly doubled in a short time owing to the dissociation:



As indicated above, about 98 per cent. of the gas dissociates in this manner. Mix the dissociated gas with sufficient oxygen to give an explosive mixture with the hydrogen, and spark the mixture. The contraction will indicate the amount of water in the gas. *E.g.*

Volume of ammonia	10.0 c.c.
Volume after sparking	19.9 c.c.
Volume after adding oxygen	72.3 c.c.
Volume after the explosion	49.9 c.c.
Contraction	22.4 c.c.

Excluding from consideration the very slight oxidation of the nitrogen, and the imperfect decomposition of the ammonia by sparking, the contraction shows that 22.4 c.c. of water has been formed, two-thirds of this, 14.9 c.c., represents the hydrogen obtained from the 10 c.c. of ammonia used for the experiment. Hence 10 c.c. of ammonia furnish very nearly 14.9 c.c. of hydrogen and 5 c.c. of nitrogen, that is, three volumes of

hydrogen per volume of nitrogen—within the limits of the experimental error.

2. By electrolysis.—If a concentrated aqueous solution of ammonia be electrolyzed (Fig. 18), nitrogen is evolved at the anode and hydrogen at the cathode. The two gases are evolved at the electrodes very nearly in the proportion of one volume of nitrogen for three volumes of hydrogen. The aqueous ammonia does not conduct very well unless a little ammonium sulphate or other ammonium salt be dissolved in the solution. A solution of ammonia in a saturated solution of ammonium nitrate furnishes a good electrolyte for this experiment.

3. By A. W. Hofmann's volumetric method (1865).—In this old experiment, a tube, Fig. 224, divided into three equal parts, is filled with chlorine gas, and concentrated aqua ammonia is run through the tap funnel A, drop by drop, until the reaction between the chlorine and the ammonia ceases. The first drop of ammonia gives a yellowish-green flame; as more ammonia is added, dense clouds of ammonium chloride are formed; much heat is evolved. Hence it is best to make the experiment with the tube immersed in a cylinder of water. When an excess of ammonia has been added, neutralize the excess with dilute sulphuric acid. The gas in the tube was originally at atmospheric pressure, it is now under reduced pressure. To restore equilibrium, a long tube is filled with dilute hydrochloric acid, and the long leg allowed to dip in a beaker of dilute hydrochloric acid while the short leg is attached to the funnel as indicated in the diagram, Fig. 224. Open the stopcock, and liquid will run into the tube until it reaches the second mark on the tube. The tube now contains one volume of nitrogen. The interpretation of Hofmann's experiment is as follows: The hydrogen of the ammonia and the chlorine combine in equal volumes to form hydrogen chloride. The hydrogen chloride combines with the ammonia to form ammonium chloride. The tube originally contained three volumes of chlorine. This chlorine has taken three volumes of hydrogen to form hydrogen chloride, etc. The latter dissolves in the liquid in the tube, and hence is without influence on the volume of the residual nitrogen. The three volumes of hydrogen were combined in ammonia with the one volume of nitrogen which remains in the tube. This proof is ingenious.

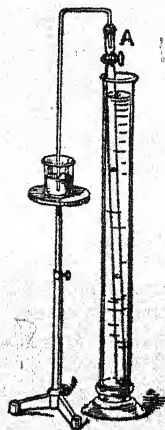
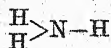


FIG. 224.—Hofmann's Apparatus.

4. The vapour density of ammonia and Avogadro's hypothesis.—The three proofs just considered ¹ show that three volumes of hydrogen combine with one volume of nitrogen to form ammonia. By Avogadro's hypothesis, the number of molecules which combine are in the same ratio; and since the two gases hydrogen and nitrogen have diatomic molecules, it follows that ammonia contains three atoms of hydrogen for one atom of nitrogen. Hence the formula of ammonia must be NH_3 , or N_2H_6 , etc. The vapour density of ammonia ($\text{H}_2 = 2$) is nearly 17. If the atomic weight of hydrogen be 1 and nitrogen 14, the molecular weight of ammonia is 17. This agrees with the number obtained for the vapour density.

¹ Others are available if needed.

Hence the formula of ammonia is NH_3 . Here the nitrogen atom appears to be tervalent, and accordingly the graphic formula for ammonia is :



5. Gravimetric analysis.—The composition of ammonia by weight can be verified by gravimetric analysis. Ammonia gas is passed over a tube containing heated copper oxide. The resulting water is weighed—Fig. 15—and the volume of nitrogen passing along is determined, and the corresponding weight computed (p. 84). The numbers so obtained give the combining proportions of hydrogen and nitrogen in ammonia. The result shows that 14.01 parts of nitrogen are combined with 3.024 parts of hydrogen. The molecular formula is then to be established by Avogadro's hypothesis.

§ 7. Molecular Compounds and Compound Salts.

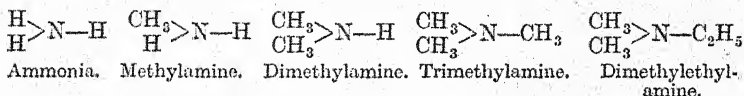
The eye may rest with complacency upon the simple beauty of the law which governs the construction of bodies belonging to the type of ammonia.—
A. W. HOFMANN.

It seems as if tervalent nitrogen in ammonia, NH_3 , changes into quinquevalent nitrogen in forming ammonium chloride, NH_4Cl . But the ready dissociation of ammonium chloride into ammonia and hydrogen chloride led A. Kekulé (1864) to assume the existence of what he designated "molecular compounds," and to reject the hypothesis of a quinquevalent nitrogen in the ammonium compounds. Kekulé adhered rigidly to the doctrine of fixed valency, and he applied the term "atomic compounds" to those compounds in which all the atoms of the elements are united in such a way that their valencies are saturated. "These," said Kekulé, "are the true chemical molecules, and the only ones which can exist in the gaseous state." He assumed that the atoms of different molecules could attract one another so as to produce a kind of coupling of the molecules. The nature of the atoms of the copulated molecules may be such that double decomposition is not possible, and "the two molecules, so to speak, adhere and form a group endowed with a certain amount of stability, which is always less than that of atomic combination." In other words, Kekulé assumed that molecular compounds are formed by direct addition of two or more simple molecules, and the simple molecules retain to some extent their individuality, for they can be readily separated from one another apparently unchanged. Kekulé cited as examples of molecular compounds: ammonium salts, phosphorus pentachloride, iodine trichloride, crystalline salts which furnish anhydrous salt and water, etc. The compounds of silver chloride and ammonia, ferrous sulphate and nitric oxide, etc., can be added to the list. To sum up Kekulé's position: in order to make the theory of constant valency compatible with the existence of more complex molecules, it was assumed that these "molecular compounds" belonged to a different type of combination. The question is not so easily answered.

If phosphorus in phosphorus trifluoride, PF_3 , be tervalent, phosphorus pentafluoride, PF_5 , must, according to Kekulé, be a molecular compound of PF_3 and F_2 ; and, as such, PF_5 should be unstable and break down into its two atomic constituents when heated. According to T. E. Thorpe

(1876), phosphorus pentafluoride is a gas stable at high temperatures. Hence the valency of an element depends on the number of atoms with which it is associated, or else the molecular compound is more stable than Kekulé supposed.

V. Meyer and Lecco (1874) applied an ingenious experiment to test whether nitrogen be ter-, or quinquivalent in the ammonium salts. It is undoubtedly tervalent in ammonia, NH_3 . Meyer and Lecco's argument will appear from the following considerations: In organic chemistry, a series of compounds is discussed in which the hydrogen atoms of ammonia are replaced, one by one, with equivalent univalent radicles—methyl, CH_3 ; ethyl, C_2H_5 ; . . .—to form a series of compounds called the amines or substituted ammonias in which nitrogen is undoubtedly tervalent. Thus:



Like ammonia, (1) the amines combine directly with acids to form compounds resembling the ammonium salts, *e.g.* trimethylamine unites directly with hydrogen chloride to form trimethylammonium chloride: $\text{N}(\text{CH}_3)_3 + \text{HCl} = \text{NH}(\text{CH}_3)_3\text{Cl}$; and (2) the amines combine with methyl or ethyl iodides, etc., to form corresponding compounds. *E.g.* trimethylamine unites with ethyl iodide: $\text{N}(\text{CH}_3)_3 + \text{C}_2\text{H}_5\text{I} = \text{N}(\text{CH}_3)_3\text{C}_2\text{H}_5\text{I}$. The properties of the compound $\text{N}(\text{CH}_3)_3\text{C}_2\text{H}_5\text{I}$ formed by the union of $\text{N}(\text{CH}_3)_3$ with $\text{C}_2\text{H}_5\text{I}$ are identical with the compound $\text{N}(\text{CH}_3)_2\text{C}_2\text{H}_5\text{CH}_3\text{I}$ formed by the union of $\text{N}(\text{CH}_3)_2\text{C}_2\text{H}_5$ with CH_3I . Consequently, the two compounds must have the same constitution. And it is inferred that the ammonium salts cannot be molecular compounds as postulated by Kekulé and that the nitrogen in ammonium compounds is not tervalent but rather quinquivalent. This is virtually Ampère's ammonium theory propounded in 1815. The argument is not quite sound, because it is possible that the groups are rearranged during the formation of the compounds by the different processes, so that the most stable configuration is always formed; and one final product is obtained by the two different reactions. This subject will be resumed when discussing Werner's theory.

Two or more simple salts, as we have seen, each primarily formed by the union of base and acid, may unite to form other salts—compound salts of greater complexity, sometimes called molecular compounds. The compound salts are often well crystallized, and they are frequently formed by replacing one or more molecules of the water of crystallization by equivalent molecules of another salt. The evidence of the state of the dissolved salt is mainly obtained from observations on the physical properties of solution—*e.g.* the rates of diffusion of the constituents; the volume changes on mixing; thermochemical phenomena; freezing points; electrical conductivities; solubilities; etc. There are three types indicated in what precedes:

1. **Mixed crystals and solid solutions.**—Potassium perchlorate and potassium permanganate are isomorphous, and form mixed crystals of all shades of colour ranging from a faint pink to a deep purple according to the relative proportions of the two salts in the crystals. The physical properties of the mixed crystals are additive, that is, continuous functions

of their compositions; and aqueous solutions of the mixed crystals give reactions characteristic of their components—in the present example, of potassium perchlorate and permanganate. The two salts form crystals containing all possible proportions between 100 per cent. potassium perchlorate and 100 per cent. potassium permanganate, as discussed under Retger's law (*q.v.*).

2. Double salts.—As in the case of mixed crystals, aqueous solutions of double salts give reactions characteristic of the component simple salts, but the physical properties of the solid salt are not necessarily additive, and the component salts only unite in simple molecular ratios. For instance, lithium chloride, LiCl , and green cupric chloride, $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$, unite to form ruby-red crystals; and an aqueous solution gives reactions characteristic of chlorides, copper, and lithium. Copper bromide also forms a dark brown aqueous solution which becomes purple-red when lithium bromide is added. The change in colour is due to the formation of a complex, $\text{Li}_2\text{CuBr}_4 \cdot 6\text{H}_2\text{O}$. Ferrous ammonium sulphate is another example, since it gives reactions characteristic of both ferrous sulphate and of ammonium sulphate. It is pure chance if the composition of a mixed crystal happens to be in simple molecular ratio since a variation in composition of the mother liquid from which the crystals are deposited will be attended by a variation in the composition of the crystals. A change in the composition of the liquid may change the composition of the compound salt, but the change will be abrupt, not gradual, in harmony with the law of multiple proportions.

3. Complex salts or salts of complex acids. As in the case of double salts, the component simple salts of the so-called complex salts are combined in a simple molecular ratio, and the resulting compound salt is quite distinct from a mechanical mixture of the component salts; and unlike double salts, the chemical properties of a solution of a complex salt are different from the properties of solutions of the component salts. The term "double salt" is often applied somewhat loosely to compound salts formed by the union of one or more molecules of one salt with one or more molecules of another salt; but, as W. Ostwald said in 1889, the term "double salt" should not be applied to combinations of two salts which give reactions different from those of the constituent salts. For example, the complex salt potassium ferrocyanide is a compound salt formed by the reaction: $4\text{KC}_y + \text{FeCy}_2 = \text{K}_4\text{FeCy}_6$. The product of this combination does not give the analytical reactions characteristic of potassium cyanide nor of ferrous cyanide. In the language of the ion theory, this is expressed by saying that the ions of complex salts in solution are different from and the ions of double salts are similar to the ions of the simple salts from which they are derived. The ions of potassium cyanide (neglecting secondary reactions) are K^+ and Cy^- ; and the ions of ferrous cyanide, Fe^{++} and Cy^- . The ions of potassium ferrocyanide, on the contrary, are K^+ and FeCy_6^{4-} . There are, therefore, no ions of Cy^- and of Fe^{++} in an aqueous solution of potassium ferrocyanide.

The physical properties of a double salt in solution may or may not be different from those of a simple mixture of the constituents, because complex salts may be completely, partially, or not at all dissociated into their constituent salts when dissolved in water. In the limiting case, the physical properties will be additively those of their components—

—complete dissociation—but in some cases, this is not the case. For instance, the solid double salt $\text{FeCl}_3 \cdot 2\text{KCl} \cdot \text{H}_2\text{O}$ is red. A concentrated aqueous solution of the double salt, or of an equivalent mixture of the component salts, at 15° has the characteristic yellow tint of ferric chloride, FeCl_3 , but at about 30° the yellow colour gives way to red. This is supposed to show that the double salt is not dissociated into its constituent molecules at 30° , but it is dissociated at 15° . Ferric chloride alone in solution does not give the red colour at 30° . Similar remarks apply to many other physical properties of double salts. A complex salt might dissociate into its constituent salts under certain conditions of temperature or concentration so that it acts as a double salt at one temperature and as a complex salt at another.

§ 8. Hydroxylamine.

Molecular weight, $\text{NH}_2\text{O} = 33.03$. Melting point, 33.05° , decomposes when heated at ordinary pressures, but boils at 58° under a pressure of 22 mm.

The preparation of hydroxylamine hydrochloride.—Hydroxylamine is formed by the reduction of nitric oxide, nitric acid, or certain nitrates. For instance, a stream of nitric oxide may be passed through a solution of tin dissolving in hydrochloric acid: $2\text{NO} + 3\text{H}_2 = 2\text{NH}_2\text{O}$. The solution will contain a compound of hydroxylamine and hydrogen chloride—hydroxylamine hydrochloride—and also tin chloride. The tin may be precipitated by passing hydrogen sulphide through the solution. The liquid is then filtered and evaporated to dryness. Extract the residue with absolute alcohol, and on evaporating the solution white crystals of hydroxylamine hydrochloride, $\text{NH}_2\text{O} \cdot \text{HCl}$, are obtained. Hydroxylamine hydrochloride is now largely made by the electrolytic reduction of nitric acid. E. Divers and T. Haga's process (1896), in which sulphur dioxide is passed into a solution containing equimolecular proportions of sodium carbonate and nitrate, is a more easily conducted process with a good yield.

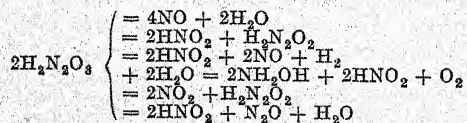
The preparation of hydroxylamine.—To isolate hydroxylamine, dissolve the hydrochloride in methyl alcohol and add sodium methylate (obtained by dissolving metallic sodium in methyl alcohol). Sodium chloride is precipitated in the alcoholic solution. Filter off the sodium chloride, and remove the alcohol by distillation—at first under ordinary pressure, and then under reduced pressure (p. 231). The hydroxylamine distils at about 70° under a pressure of 60 mm.; or at, say, 58° under a pressure of 22 mm. When distilling hydroxylamine, provision must be made for an explosion.

Properties.—Hydroxylamine crystallizes in the form of white needle-like crystals without smell. The crystals melt at 33° , and boil at 58° under a pressure of 22 mm. At ordinary pressures, hydroxylamine gradually decomposes if heated over 15° , and at higher temperatures it is liable to decompose explosively with a yellow flash. Hydroxylamine resembles ammonia in many respects—it dissolves in water, forming a strongly alkaline solution; it reacts with acids, forming salts— $\text{NH}_2\text{O} \cdot \text{HCl}$; $(\text{NH}_2\text{O})_2 \cdot \text{N}_2\text{SO}_4$; $\text{NH}_2\text{O} \cdot \text{HNO}_3$, etc. Hydroxylamine hydride, NH_2OH . HI, for instance, is made by mixing molecular quantities of hydroxylamine

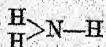
and hydrogen iodide in aqueous solutions, and evaporating at about 20° in vacuo. It forms colourless, hygroscopic, acicular crystals. The salts all decompose more or less violently when heated. The nitrate furnishes nitric oxide and water; and the nitrite furnishes nitrous oxide and water. Compounds like NH_2ONa and $(\text{NH}_2\text{O})_2\text{Ca}$ are known as hydroxylamites. Hydroxylamine thus behaves towards alkalis like a weak acid, and towards acids like a weak base, *i.e.* hydroxylamine behaves as if it were both an acid and a base.

Hydroxylamine salts are used as reducing agents in analytical work. Hydroxylamine precipitates metallic silver from silver nitrate; it reduces mercuric chloride to mercurous chloride; it precipitates cuprous oxide from cupric salts; chromium hydroxide from chromic salts, etc. It oxidizes greenish ferrous hydroxide, suspended in an alkaline solution, to red ferric hydroxide, and at the same time ammonia is formed. If the precipitate be dissolved in an acid, and treated with hydroxylamine, the solution becomes colourless owing to the reduction of the red ferric to greenish ferrous salt.

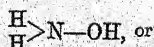
If a solution of hydroxylamine and sodium hydroxide in methyl alcohol be treated with methyl nitrate, $\text{CH}_3\cdot\text{NO}_3$, a white powder is precipitated. It has the empirical composition $\text{Na}_2\text{N}_2\text{O}_3$. This substance appears to be the sodium salt of a very unstable acid—nitrohydroxylamic acid, $\text{H}_2\text{N}_2\text{O}_3$ —for if the salt be treated with a mineral acid, the product which is liberated begins at once to decompose with brisk effervescence, forming nitric oxide and water: $\text{H}_2\text{N}_2\text{O}_3 \rightarrow 2\text{NO} + \text{H}_2\text{O}$. The acid thus appears to be a hydrate of nitric oxide. In addition to the sodium salt, some organic derivatives have been made. The constitutional formula is usually written $\text{HO}-\text{N}=\text{NO}-\text{OH}$. The sodium salt rapidly absorbs oxygen from the air, forming sodium nitrite and nitrate, and when the solution is boiled it furnishes sodium nitrite, etc., as symbolized in the equation: $2\text{Na}_2\text{N}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow 2\text{NaNO}_2 + \text{N}_2\text{O} + 2\text{NaOH}$; the heating of the solid salt furnishes sodium nitrite and hyponitrite: $2\text{Na}_2\text{N}_2\text{O}_3 = 2\text{NaNO}_2 + \text{Na}_2\text{N}_2\text{O}_2$. The acid and its salts are interesting in that they may decompose in at least six different ways, as symbolized:



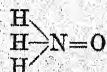
Composition.—The molecular weight by the freezing-point process corresponds with NH_3O . It is generally supposed that hydroxylamine has a similar constitution to ammonia, but one hydrogen atom of the ammonia is replaced by hydroxyl:



Ammonia.



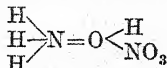
Hydroxylamine.



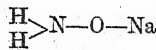
Hydroxylamine.

Hence the alternative term "oxyammonia" for hydroxylamine. Some consider that the graphic formula should be written $\text{H}_3\equiv\text{N}=\text{O}$; but the compound seems to exhibit a kind of desmotropism, forming what are

called oxonium salts when it acts as a base (oxygen quadrivalent), and hydroxylamates when it acts as an acid :



Hydroxylamine nitrate.



Sodium hydroxylamate.

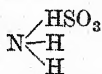
History.—Hydroxylamine salts, and the aqueous solution of the base, were first prepared by W. Lossen (1865); and anhydrous hydroxylamine was simultaneously made by different processes by L. de Bruyn and by L. Crismer in 1890–91.

§ 9. Some Derivatives of Ammonia and Hydroxylamine.

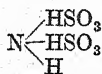
The group NH_2 behaves as a monad radicle, and compounds containing it are called amides or amines. Do not confuse this term with the so-called addition compounds of ammonia also called **ammines**—*vide* Chap. XXXVI. § 5. Compounds with the dyad group NH are called *imides*. By the action of sulphur trioxide on ammonia, or by saturating hydroxylamine hydrochloride with sulphur dioxide, **amidosulphonic acid**, NH_2HSO_3 , separates in colourless rhombic prisms. The ammonium salt of imido-sulphonic acid, $\text{NH}(\text{HSO}_3)_2$, is formed by the action of ammonia on chloro-sulphonic acid; and **nitrilsulphonic acid**, $\text{N}(\text{HSO}_3)_3$, is formed by the prolonged action of sulphurous acid on nitrous acid. The relations of these compounds to ammonia are illustrated by the scheme :



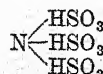
Ammonia.



Amidosulphonic acid.

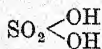


Imidosulphonic acid.

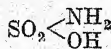


Nitrilsulphonic acid.

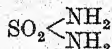
Amidosulphonic acid may be regarded as sulphuric acid with one of its hydroxyl-groups replaced by an amide-group, $\text{NH}_2\text{—SO}_2\text{—OH}$, and in sulphamide, $\text{SO}_2(\text{NH}_2)_2$, both the hydroxyl groups of sulphuric acid are replaced by the amide group; while in sulphimide, SO_2NH , the two hydroxyl groups are replaced by the dyad radicle NH .



Sulphuric acid.



Amidosulphuric acid.

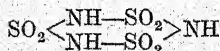


Sulphamide.



Sulphimide.

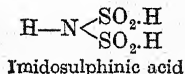
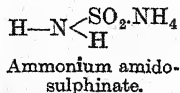
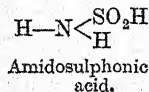
The molecular weight of sulphimide corresponds with the polymer $(\text{SO}_2\text{NH})_3$, and it is hence called **trisulphimide**,



When a solution of sulphuryl chloride in an indifferent solvent—say chloroform—is saturated with ammonia, and the product of the action is extracted with water and shaken with lead or silver oxide to remove the chlorine, the clear solution on evaporation furnishes sulphamide. This substance forms compounds with metallic oxides, and when heated to about 200° it is transformed into sulphimide: $\text{SO}_2(\text{NH}_2)_2 = \text{NH}_3 + \text{SO}_2\text{NH}$. Several salts of sulphimide are known—*e.g.* $\text{SO}_2\text{N.Na}$, $\text{SO}_2\text{N.Ag}$, $(\text{SO}_2\text{N})_2\text{Ba}$, etc. When sulphuryl chloride and liquid ammonia react

at a low temperature, ammonium imidosulphamide, $\text{H.N}(\text{SO}_2.\text{NH}_2)_2$, is formed: $2\text{SO}_2\text{Cl}_2 + 7\text{NH}_3 = 4\text{NH}_4\text{Cl} + \text{NH}(\text{SO}_2.\text{NH}_2)_2$. When the mixture is acidified with hydrochloric acid and warmed, ammonium chloride, sulphuric acid, and sulphamide are produced. The latter alone is extracted by dry acetic ether. Sulphamide melts at 81° , and trisulphimide at 165° .

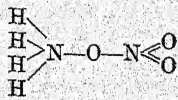
The final product of the interaction between nitrites and sulphites, or between nitrous and sulphurous acid is the complete oxidation of the latter to sulphuric acid, but quite a number of acids, or rather salts of acids, representing intermediate stages in the reaction have been separated. These intermediate products are generally regarded as hydroxylamine-sulphonic acids. Thus, when an excess of sulphur dioxide acts on ammonia, amidosulphonic acid, $\text{NH}_2.\text{HSO}_3$, is formed; and with an excess of ammonia, either the white ammonium salts of amidosulphonic acid, or a red-coloured isomer, $2\text{SO}_2.4\text{NH}_3$, or probably the ammonium salt of imidosulphonic acid, $(\text{HSO}_2)_2\text{NH}$, is formed.



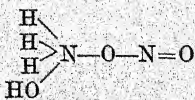
§ 10. Isomerism.

The great interest of isomerism has been to introduce into science the principle that substances may be, and are, essentially different solely because the arrangement of the atoms in their chemical molecules is not the same.—L. PASTEUR.

By an illogical application of the law of constant composition it was once thought to be self-evident "that substances which contain the same atoms and the same relative quantities of these must of necessity have the same chemical properties." True enough, composition is the most fundamental of all chemical properties, but constitution also is of paramount importance, we now believe that the atoms of a molecule are arranged according to plan so definite and precise that two different atoms cannot change places without altering the properties of the substance. This does not mean that the relations of the atoms for one another are necessarily immovable, for, as indicated on p. 146, the atoms may revolve about a position of equilibrium without altering their order of succession. Ammonium nitrate and hydroxylamine nitrite are two different substances with the same ultimate composition, the same molecular weight, and both furnish nitrous oxide and water when heated. There the similarity almost ends. The general properties of the two salts are so very different that there is little room for doubt that the constitution of the molecules must be quite different. There are at least two modes in which the atoms may be arranged for stable systems. The probable constitutional formulæ of the two compounds are:



Ammonium nitrate.



Hydroxylamine nitrite.

We have met several compounds which have the same ultimate composition, but a different molecular weight, and different properties. *E.g.* NO_2 and N_2O_4 , the α - and β -sulphur trioxides, etc. It is convenient to fix these ideas definitely by the use of the special term isomerism—from the Greek *isos* (isos), the same; *μέρος* (meros), part. Isomerism is a general term applied when the percentage composition of two or more substances is the same, but the properties are different. The term "allotropism" is reserved for the special case of isomerism among elements, p. 483, and "isomerism" for compounds. Polymerism is applied when the percentage composition of two or more substances is the same, but the molecular weight is different, p. 482. It is *supposed* that in most cases allotropism is a special case of polymerism. Metamerism—from the Greek *μετά* (meta), change—is applied when the percentage composition and molecular weights are the same, but the properties are different—*e.g.* hydroxylamine nitrite and ammonium nitrate. Desmotropism or tautomerism, p. 510, and physical, geometrical or enantiomorphical isomerism—the stereoisomerism of the tartaric acids—discussed on p. 628, are special cases of metamerism. To summarize:¹

Same ultimate composition; properties different:	
I. Elements	ALLOTRIPISM
II. Compounds	ISOMERISM
(1) Molecular weights different	POLYMERISM
(2) Molecular weights the same	METAMERISM
(a) Relative positions of some radicles labile, not fixed	DESMOTROPISM
(b) Certain radicles are enantiomorphical	GEOMETRICAL ISOMERISM

§ 11. Hydrazine or Diamide.

Molecular weight, $\text{N}_2\text{H}_4 = 32.05$. Melting point, 1.4° ; boiling point, 113.5° ; critical temperature, 380° .

It is possible to make ammonia double on itself, so to speak, by converting one part into chloramine, NH_2Cl , and bringing this in contact with more ammonia: $\text{NH}_3 + \text{NH}_2\text{Cl} = \text{NH}_2\text{NH}_2\text{HCl}$. This constructive reaction is opposed by a destructive change: $2\text{NH}_3 + \text{NH}_2\text{Cl} = 3\text{NH}_4\text{Cl} + \text{N}_2$, which is considerably hampered if the viscosity of the solution be increased by adding glue; and it is favoured by the addition of acetone.

The preparation of hydrazine sulphate.—Several convenient methods of preparing hydrazine are described in text-books of organic chemistry, but it may be obtained quite as conveniently from purely inorganic substances. Mix 200 c.c. of a 20 per cent. solution of ammonia, 5 c.c. of a one per cent. solution of glue or gelatine,² and 100 c.c. of an aqueous solution of 7.5 grams of sodium hypochlorite—free from an excess of chlorine—in a litre flask. Boil the mixture for about half an hour, when it will have evaporated to about half its original volume.

¹ In different text-books there is some difference as to the exact meaning assigned to the different terms here summarized. There are also other peculiarities not covered by the terms in the text.

² The action of the gelatine is not understood. The glue and gelatine are not necessary, but the yield of hydrazine is much reduced if the mucilage be absent. This is an empirical fact. It is generally stated that gelatine prevents the formation of nitrogen chloride.

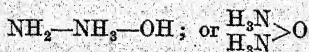
Monochloramide is first formed (p. 655), and this reacts with another molecule of ammonia to form hydrazine hydrochloride: $\text{NH}_2\text{Cl} + \text{NH}_3 = \text{N}_2\text{H}_4\cdot\text{HCl}$. When cold, place the flask in iced water, and add 20 c.c. of a solution containing 1.96 grams of sulphuric acid, H_2SO_4 . Hydrazine sulphate crystallizes out. This may be purified by recrystallization from water. The process is used technically for the preparation of hydrazine sulphate.

The preparation of hydrazine hydrate and hydrazine.—When hydrazine sulphate is distilled with potassium hydroxide in a silver vessel, screwed at the junctions, hydrazine hydrate is obtained: $\text{N}_2\text{H}_4\cdot\text{H}_2\text{SO}_4 + 2\text{KOH} = \text{N}_2\text{H}_4\cdot\text{H}_2\text{O} + \text{K}_2\text{SO}_4 + \text{H}_2\text{O}$. The free base hydrazine is made by adding small quantities of the hydrate to barium oxide in a glass flask cooled in a freezing mixture. The mixture is then distilled under reduced pressure. The barium oxide removes the water from the hydrate: $\text{BaO} + \text{N}_2\text{H}_4\cdot\text{H}_2\text{O} = \text{Ba}(\text{OH})_2 + \text{N}_2\text{H}_4$. F. Raschig's method of distilling with sodium hydroxide gives a better yield.

Properties.—Hydrazine hydrate is a colourless fuming corrosive liquid soluble in water. It boils at 118° , and freezes to a white crystalline solid, melting at -40° . It attacks glass, cork, and rubber. It is strongly basic, and forms a series of salts with the acids, e.g. hydrazine monochloride, $\text{N}_2\text{H}_4\cdot\text{HCl}$; and hydrazine dichloride, $\text{N}_2\text{H}_4\cdot 2\text{HCl}$. Hence hydrazine is a diacid base. Most of the salts are very soluble in water. The sulphate, $\text{N}_2\text{H}_4\cdot\text{H}_2\text{SO}_4$, is not so soluble. Hydrazine and its salts are among the most powerful reducing agents known; they reduce cupric salts to red cuprous oxide, and precipitate metallic silver from silver nitrate, mercury from mercuric chloride, etc. Hydrazine and sulphur dioxide react to form the hydrazine salt of an unknown hydrazine disulphonic acid, $\text{HSO}_2\cdot\text{NH}\cdot\text{NH}\cdot\text{HSO}_2$, as indicated in the equation: $2\text{SO}_2 + 9\text{N}_2\text{H}_4 = 3\text{N}_2\text{H}_2(\text{HSO}_4\cdot\text{N}_2\text{H}_4)_2$. Thionyl chloride is reduced to sulphur which dissolves in the excess of hydrazine furnishing sulphohydrazinium. Sulphur trioxide produces sulphur sesquioxide.

The free base is a colourless, fuming liquid boiling at 56° under a pressure of 71 mm., and at 113.5° under a pressure of 761 mm. The liquid freezes at 0° , and melts at 1.4° . When heated to about 350° it decomposes into ammonia and nitrogen: $3\text{N}_2\text{H}_4 = \text{N}_2 + 4\text{NH}_3$. Hydrazine is slowly oxidized in air with the liberation of free nitrogen, and it burns with a violet-coloured flame.

Composition.—The vapour density of the hydrate at 100° is 50. This corresponds with the molecule $\text{N}_2\text{H}_4\cdot\text{H}_2\text{O}$. The hydrate dissociates into water and free base *in vacuo* at 100° , and at 143° dissociation is complete. At 183° , under atmospheric pressure, the hydrate decomposes into nitrogen and ammonia. A freezing-point determination of the aqueous solution gives a molecular weight corresponding with the dihydrate: $\text{N}_2\text{H}_4\cdot 2\text{H}_2\text{O}$. The constitution of the hydrate is either



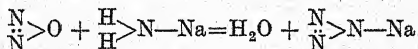
History.—Hydrazine hydrate was discovered by T. Curtius in 1887, and the base was isolated by L. de Bruyn, 1895. The substitution products—e.g. phenylhydrazine, $\text{C}_6\text{H}_5\cdot\text{HN}\cdot\text{NH}_2$ —have been known for a longer time. E. Fischer made the first organic derivative in 1875.

§ 12. Hydrazoic Acid, Hydronitric Acid, or Azoimide.

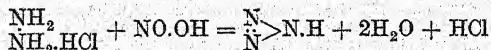
Molecular weight, $N_3H = 43.04$. Melting point, -80° ; boiling point, 37° .

The successful attempt to double ammonia on itself, resulting in the discovery of hydrazine, was followed by a further attempt to replace another hydrogen atom by a NH_2 group, and so form triamine or triazine $H_2N.NH.NH_2$; but this could not be accomplished, for the terminal hydrogen atoms dropped out and a closed ring of three nitrogen atoms, N_3H , was the result. This so-called triazo-group forms a large class of compounds many of which are artificial colouring agents or dyes.

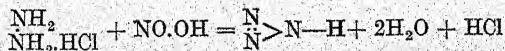
Preparation.—Like hydrazine, this compound is best made by organic processes. W. Wislicenus (1892) made it from inorganic materials by passing dry ammonia over metallic sodium in a nickel boat in a tube between 250° and 350° so as to make sodamide: $2Na + 2NH_3 = 2NaNH_2 + H_2$. When all the sodium has been converted into the amide (about six hours), the current of ammonia is replaced by a stream of dry nitrous oxide, and continued at 190° until ammonia is no longer evolved (about five hours):



The product of the action—a mixture of $NaOH$ and NaN_3 —is dissolved in water; the solution acidified with dilute sulphuric acid (1 : 1); and distilled. The first quarter of the distillate contains most of the hydrazoic acid. Hydrazoic acid can also be made by treating nitrogen chloride, NCl_3 , or an aqueous solution of silver nitrite with hydrazine sulphate. In the latter case a crystalline precipitate of silver hydrazoate, N_3Ag , is formed in a short time—A. Angeli (1893). The reaction may be symbolized:



When nitrous acid acts on ammonia, both atoms of nitrogen are liberated in the elemental form, $NH_3 + HNO_2 = 2H_2O + N_2$; but T. Curtius found in 1890 that when hydrazine is used in place of ammonia, the attack by nitrous acid furnishes the simplest known form of triazo-group, namely, hydrazoic acid, N_3H . The isolation of this acid, said D. I. Mendeleeff, is one of the most important achievements of 1890. The reaction is symbolized: $HNO_2 + N_2H_4 = HN_3 + 2H_2O$. Instead of using free nitrous acid, Stollé and Thiele obtained better results with ethyl or amyl nitrite in an alkaline solution, say, $N_2H_4 + C_2H_5NO_2 + NaOH = NaN_3 + C_2H_5OH + H_2O$. The resulting sodium azide is a commercial product. The preparation of hydrazoic acid by the action of nitrous acid on salts of hydrazine resembles a well-known method of preparing nitrogen from ammonium salts by the action of nitrous acid—say by heating ammonium chloride with sodium nitrite—writing nitrous acid in place of the latter:

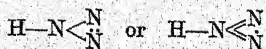


The aqueous solution of the free acid can be obtained by distillation with sulphuric acid as indicated above; repeated fractional distillation furnishes a solution containing 91 per cent. of hydrazoic acid. The remaining water must be removed by calcium chloride.

Properties.—Pure hydrazoic acid is a colourless mobile liquid, with an unpleasant penetrating odour. Trifling quantities produce headaches, and the vapour is poisonous. The acid boils at 37° ; it can be solidified by cooling. The solid melts at -80° . Work with hydrazoic acid and its salts is dangerous because these compounds are rather unstable, and very liable to explode. The acid is partly reduced by sodium amalgam, forming ammonia and a little hydrazine: $\text{N}_3\text{H} + 3\text{H}_2 = \text{N}_2\text{H}_4 + \text{NH}_3$. Potassium permanganate oxidizes it to water and nitrogen; oxygen is also liberated at the same time. When hydrazine sulphate is treated with ammonium metavanadate and sulphuric acid, some ammonia, NH_3 , and hydrazoic acid are formed: $2\text{N}_2\text{H}_4 + \text{O}_2 = \text{NH}_3 + \text{HN}_3 + 2\text{H}_2\text{O}$. The experimental yield of hydrazoic acid is but 13.6 per cent. of the theoretical.

The acid is soluble in water. The aqueous solution behaves as a strong monobasic acid, and it readily dissolves zinc, iron, cadmium, magnesium, and aluminium with the evolution of hydrogen and ammonia, and the formation of salts—called hydrazoates, or azides, or trinitrides. It also gives insoluble silver and mercury salts, AgN_3 and HgN_3 , both of which are very explosive; the silver, lead, mercurous, and cuprous azides are very sensitive to light, and decompose directly into metal and nitrogen. The ferric salt has an intense red colour, and the colour can be detected when a ferric salt is added to a solution containing one part of hydrazoic acid per million. One part of the acid per million will give a perceptible turbidity with silver nitrate. Lead azide, PbN_6 , is used as a detonator in place of mercury fulminate for trinitrotoluol; it is considered less sensitive and safer. The salts are usually anhydrous and crystalline, and when heated give the pure metal. The aqueous solution and the alkaline salts are not so liable to explode as the salts of the heavy metals; in many ways it behaves chemically like hydrochloric acid, with ammonia it yields the ammonium salt: $\text{NH}_3 \cdot \text{HN}_3$, or NH_4N_3 , that is, N_4H_4 ; and with hydrazine, $\text{N}_2\text{H}_4 \cdot \text{HN}_3$, that is, N_5H_5 . With sodium hypochlorite and acetic acid, hydrazoic acid furnishes a colourless highly explosive compound, chlorazide, N_3Cl , which smells like hypochlorous acid, and is decomposed by an aqueous solution of sodium hydroxide, $\text{N}_3\text{Cl} + 2\text{NaOH} \rightleftharpoons \text{NaN}_3 + \text{NaOCl} + \text{H}_2\text{O}$. Silver azide with a cold ethereal solution of iodine, furnishes a yellow explosive compound iodozide, N_3I , thus, $\text{N}_3\text{Ag} + \text{I}_2 = \text{AgI} + \text{N}_3\text{I}$.

Constitution.—On electrolysis, hydrazoic acid yields rather less than three volumes of nitrogen per one volume of hydrogen. The low yield of nitrogen is due to secondary reactions. Analyses and vapour density determinations agree with the formula N_3H . The formation of hydrazoic acid from sodamide and nitrous oxide, which results in the formation of nitrogen and ammonia, agree with the structural formula:



if one of the nitrogen atoms be quinquevalent. The resemblances between this acid and hydrochloric acid suggests that the triazo-group, N_3 , in its chemical properties bears a strong family resemblance to the halogens. The physical properties—light refraction, light dispersion, boiling points,

electrical conductivities, etc.—of compounds of the organic triazo-derivatives and the corresponding halogen compounds are very similar. Thus :

	Acetic acid	Iodo-	Triazo-	Bromo-	Chloro-acetic acid.
Ionization constant $\times 10^4$	0.018	7.5	9.3	13.8	15.5

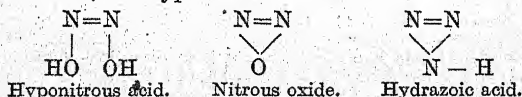
With metals below magnesium in the electrochemical series, Table XXIII., hydrazoic acid is reduced to ammonia, and maybe to hydrazine and free nitrogen ; no hydrogen is evolved, *e.g.* $\text{Cu} + 3\text{HN}_3 = \text{CuN}_6 + \text{N}_2 + \text{NH}_3$; with nitric acid, it will be remembered that $3\text{Cu} + 8\text{HNO}_3 = 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$. With hydrides of some of the oxidizable non-metals, the acid is reduced, forming ammonia and nitrogen, thus $\text{H}_2\text{S} + \text{HN}_3 = \text{S} + \text{N}_2 + \text{NH}_3$; with nitric acid, we have $3\text{H}_2\text{S} + 2\text{HNO}_3 = 3\text{S} + 2\text{NO} + 4\text{H}_2\text{O}$. With metals lying near the end of the electrochemical series, the acid is reduced, forming ammonia and nitrogen, $\text{Pt} + 2\text{HN}_3 + 4\text{HCl} = \text{PtCl}_4 + 2\text{N}_2 + 2\text{H}_3\text{N}$; with nitric acid under similar conditions, $3\text{Pt} + 4\text{HNO}_3 + 12\text{HCl} = 3\text{PtCl}_4 + 4\text{NO} + 8\text{H}_2\text{O}$.

Assuming that if two compounds behave in a similar manner they have an analogous structure, it follows if nitric acid has the structure

$\text{H}-\text{O}-\text{N} \begin{smallmatrix} \nearrow \text{O} \\ \searrow \end{smallmatrix}$ with a quinquevalent nitrogen atom as a nucleus united with

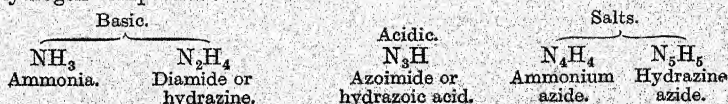
oxygen ; then in hydrazoic acid we can expect a quinquevalent nitrogen united with nitrogen, or $\text{H}-\text{N}=\text{N} \equiv \text{N}$ (J. W. Turrentine, 1912) ; hence the alternative term, "hydronitric acid," for this compound. The latter formula also agrees with some reactions studied in organic chemistry (Thiele, 1911).

The following graphic formulæ show the relations between hydrazoic acid, nitrous oxide, and hyponitrous acid :



There are many resemblances between this acid and hydrochloric acid ; which suggests that the triazo-group, N_3 , in its chemical properties bears a strong family resemblance to the halogens. Attempts have been made to prepare a hexatomic-nitrogen group, N_3N_3 , by making hydrazoic acid double on itself.

Nitrogen gas is rivalled only by argon and its congeners in its reluctance to take part in chemical changes, but when united with other elements it leads to forms of activity in great profusion, so much so that M. O. Forster (1911) styled nitrogen the most versatile form of elemental matter, for union with different proportions of hydrogen furnishes three highly active substances—ammonia, hydrazine, and hydrazoic acid—which have characteristics which stand in marked contrast with one another. If oxygen be brought into the system, hydroxylamine, nitrous acid, and nitric acid may be mentioned as typical materials capable of entering into chemical changes of the most diverse order. Enumerating the nitrogen-hydrogen compounds :



To these can be added hydrazine hydrazone, N_2H_4 , and possibly also basic ammonium, NH_4 ; and *diimide*, N_2H_2 , which is said to have been made by heating ammonium chloroplatinate: $(NH_4)_2PtCl_6 = Pt + 6HCl + N_2H_2$; but there is some doubt if this is really the case.

Questions.

1. Calculate the vapour density of ammonium chloride. By experiment it is found to be 13.345. How do you explain the difference between the calculated and the observed results? Can you give any experimental evidence in support of your explanation? Do you know of any other similar cases?—*Science and Art Dept.*

2. Calculate the weight of nitrogen contained in one cwt. of (a) ammonium sulphate; (b) sodium nitrate. Describe exactly how you would distinguish these two substances when mixed together in solution.—*Aberdeen Univ.*

3. How would you prepare in the laboratory a vessel full of ammonia? Describe experiments by which you could demonstrate that ammonia (a) is very soluble in water, (b) combines with acids to form salts, (c) contains hydrogen.—*Victoria Univ., Manchester.*

4. What is the action of ammonia gas on (a) hot copper oxide, (b) hydrochloric acid, (c) chlorine water?—*St. Andrews Univ.*

5. Why are the compounds formed by the union of acids with ammonia termed ammonium compounds? Quote facts which may be regarded as evidence that a solution of ammonia in water contains ammonium hydroxide.—*London Univ.*

6. What are the two main sources of supply of inorganic nitrogen compounds? Outline three methods which have been suggested for the production of ammonia from atmospheric nitrogen.—*Sheffield Univ.*

7. How can it be shown that the vapour obtained by heating slightly moist ammonium chloride consists of a mixture of ammonia and hydrogen chloride gases? What is the density relative to hydrogen of the vapour given off by ammonium chloride; and what would be the density if the vapour consisted of ammonium chloride? ($N = 14$; $Cl = 35.5$.) Mention other decompositions of a similar nature with which you are acquainted.—*Univ. North Wales.*

8. Give the volume relations between the gases in the following reactions: Hydrogen and chlorine combining to form hydrochloric acid, hydrogen and oxygen combining to form steam, hydrogen and nitrogen combining to form ammonia. If a million molecules of hydrogen took part in each reaction, how many molecules of each product would be formed?—*Sheffield Scientific School, U.S.A.*

9. Explain why the formula of ammonia gas is given as NH_3 , and give an account of any theories involved in your answer.—*Aberystwyth Univ.*

10. Define the terms "acid," "base," "acid salt," "basic salt," "double salt," "complex salt," and supply one illustration of each.—*Sheffield Univ.*

11. What is hydroxylamine? Describe and explain the process by which the hydrochloride is prepared from potassium nitrite. What is its action upon a solution of a cupric salt; and to what useful purpose has hydroxylamine been applied?—*Science and Art Dept.*

12. What is the effect of passing chlorine gas through aqueous solutions of (a) ammonia, (b) sulphur dioxide, (c) ferrous sulphate? What is the effect of passing sulphur dioxide gas through (a) aqueous solutions of potassium permanganate, (b) concentrated nitric acid? Give equations.—*St. Andrews Univ.*

13. How is ammonia gas prepared pure and dry? Give the names and formulae of the compounds derived from ammonia by replacing one or more of its atoms of hydrogen by (a) hydroxyl, (b) chlorine, (c) methyl. Describe the preparation and properties of any two of these compounds.—*London Univ.*

14. What change takes place when chlorine is passed into a solution of ammonia? How can the experiment be conducted so as to afford evidence of the constitution of ammonia?—*Board of Educ.*

15. How is the so-called ammonium amalgam prepared? Describe its principal properties and explain the various views which have been held as to its nature.—*Board of Educ.*

16. By what means has the relative rate of diffusion of gases been determined?

Describe an experiment by which you might test whether a gas was dissociated at a high temperature into simpler constituents. Two vessels *A* and *B* containing respectively chlorine, and a mixture of an inert gas with 10 per cent. of oxygen, were put into communication through a small hole. After diffusion had taken place for a short time, the chlorine in *A* was absorbed by potash, and the residual gas in *A* was found to contain 11 per cent. of oxygen. What was density, approximately, of the inert gas?—*Victoria Univ., Manchester.*

17. Plot the specific gravities of aqueous solutions of ammonia with their concentrations from the following data :

Grams of NH_3 per litre	308.3	255.0	210.9	157.4	95.1	4.5
Specific gravity. . .	0.882	0.900	0.916	0.936	0.960	0.998

Show (a) the weight of NH_3 in a litre of aqua ammonia of specific gravity 0.890, and (b) the per cent. of NH_3 in a solution of specific gravity 0.882. Answers: (a) 282.6, (b) 34.95.

18. Two tubes, provided with platinum wires and standing over mercury, contain severally ammonia and the products of the decomposition of ammonia; into the latter a few drops of dilute sulphuric acid are passed. A stream of electric sparks is sent through the two tubes till no further change occurs. Describe and explain the changes which have occurred.—*New Zealand Univ.*

19. Give examples of substances with "abnormal vapour densities." Explain fully the experimental evidence why the term is applied to the substances you mention, and justify its application.—*Cape Univ.*

20. Give an account of the compounds containing only nitrogen and hydrogen.—*Trinity Coll.*

21. Give the formulæ, modes of preparation, and chief properties of the following mercurial compounds: Red precipitate, white precipitate, corrosive sublimate, calomel, vermilion, and cinnabar.—*London Univ.*

22. Sal ammoniac may be represented by the symbols $\text{NH}_3\cdot\text{HCl}$ or NH_4Cl . State the evidence that can be adduced in support of each of these formulæ.—*London Univ.*

23. What experiment would you make in order to show that nitric acid may be reduced to ammonia? Describe precisely how you would identify the ammonia.—*London Univ.*

CHAPTER XXIX

NITROGEN AND ATMOSPHERIC AIR

§ 1. Nitrogen—Occurrence and Preparation.

Atomic weight, $N = 14.01$; molecular weight, $N_2 = 28.02$. Bi-, ter-, and quinque-valent. Melting point, -210.5° ; boiling point, -195.5° ; critical temperature, -146° . Relative vapour density ($H_2 = 2$), 27.81; (air = 1), 0.967. One litre under normal conditions weighs 1.2506 grams.

History.—It is difficult to state precisely who first isolated nitrogen and clearly recognized it as a definite substance. John Mayow (1674) and several others got very near, if they did not get actually there. The man who deduces on good mental evidence, or even proves by actual experiment, the existence of something not known before is not always recognized as the discoverer; but rather is he hailed discoverer who proves by a conclusive series of experiments that the substance in question has properties distinct from all other substances. *He only discovers who proves.* Otherwise, Boyle or Paracelsus would be called the discoverer of hydrogen; Lucretius of carbon dioxide; J. Kunckel of ammonia; Eck de Sulzbach of oxygen; Hooke or Cassini of the law of gravitation; etc. We must remember, said I. Todhunter, that he alone is the true discoverer who reveals, not in dim oracles which enthusiastic votaries may subsequently interpret into truth, but in clear characters which contemporaries, even though hostile, cannot misunderstand or misinterpret.

D. Rutherford (1772) is generally credited with the discovery of nitrogen. He published a thesis in Latin in 1772 in which he said:

By the respiration of animals, healthy air is not merely rendered mephitic (that is, charged with carbon dioxide), but it also suffers another change, for, after the mephitic portion is absorbed by a solution of caustic alkali, the remaining portion is not rendered salubrious, and, although it occasions no precipitate in lime-water, it nevertheless extinguishes flame, and destroys life.

Rutherford removed oxygen from the air by such combustibles as phosphorus, charcoal, etc., and washed out the products of combustion by alkalies or lime water. The residue was called by him "phlogisticated air."¹ H. Cavendish confirmed this experiment in 1785. Lavoisier first called the residue "mephitic air," and afterwards "azote." J. A. C. Chaptal (1823) suggested the name *nitrogène* from the Greek *νίτρον* (nitron), saltpetre; and *γεννάω* (gennao), I produce—because the gas is a constituent of nitre.

Occurrence.—Nitrogen constitutes four-fifths of the total volume of atmospheric air. According to spectroscopic observations it is probable that certain nebulae contain nitrogen. It is also found in certain minerals,

¹ Hydrogen was also called "phlogisticated air."

where it is probably occluded or adsorbed. It occurs combined in ammonia, nitre, and a great many animal and vegetable products—*e.g.* white of egg, proteids, etc. It is a constant and essential constituent of all living organisms; all life seems to depend upon the transformation of proteid compounds.

Preparation.—Nitrogen is easily obtained from air by removing the admixed carbon dioxide and oxygen. This is conveniently done by burning a piece of phosphorus in a dry crucible floating on the surface of water under a small bell-jar. The phosphorus combines with most of the oxygen, forming phosphorus pentoxide, and this quickly dissolves in water, leaving behind the nitrogen. If the water be alkaline with sodium hydroxide, the carbon dioxide, normally present in air, will also be removed. The nitrogen so prepared is not pure because the phosphorus ceases to burn before all the oxygen has been removed. A solution of cuprous chloride (*p.* 292) in hydrochloric acid rapidly absorbs oxygen from air, and leaves behind the nitrogen. It is best to remove the carbon dioxide by first

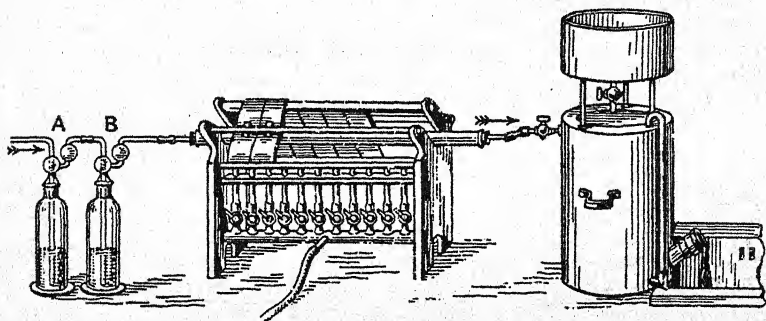


FIG. 225.—Preparation of Nitrogen.

passing the air through a solution of sodium hydroxide; and to absorb the oxygen by means of an element which will form a non-volatile oxide. Copper turnings are generally considered best for the purpose; the "turnings" offer a large surface of oxidizable metal to the air. The process is as follows:—

Air freed from carbon dioxide in a wash-bottle of sodium hydroxide, *A*, Fig. 225, and from moisture by passage through sulphuric acid, *B*, is then passed through a red-hot tube containing copper turnings. The copper removes the oxygen and forms cupric oxide: $2\text{Cu} + \text{O}_2 = 2\text{CuO}$. The nitrogen passes on to be collected in a gas jar, or gasholder, etc. In the diagram, the air is supposed to be drawn over the copper, the gas-holder being filled with nitrogen. If the gas-holder were placed at the end *A*, and air forced along the tubes, the nitrogen gas could be collected in gas jars, Fig. 225. The process of oxidation of course ceases when all the copper is oxidized. If the wash-bottle, *B*, of concentrated sulphuric acid be replaced by an aqueous solution of ammonia, as recommended by S. Lupton (1876), the ammonia reduces the copper oxide as fast as it is formed: $\text{CuO} + 2\text{NH}_3 = \text{Cu} + 3\text{H}_2\text{O} + \text{N}_2$. Any excess of ammonia can be removed by passing the gas from the copper tube through a solution

of sulphuric acid before it is collected in the gasholder. Cold boiled water should be used in the gasholder so as to lessen the risk of contamination owing to the presence of oxygen dissolved in ordinary water.

We shall see later that "atmospheric nitrogen" contains argon. Nitrogen called "chemical nitrogen" can be prepared free from argon by heating a concentrated solution of ammonium nitrite in a glass flask or retort: $\text{NH}_4\text{NO}_2 = 2\text{H}_2\text{O} + \text{N}_2$. The reaction need only be just started by gently warming the flask, it will then continue, and maybe increase in velocity without a further application of heat because the system itself becomes warmer owing to the degradation of energy. If heat be continuously applied to the flask, the reaction may become uncontrollably violent. Instead of using ammonium nitrite, a mixture of ammonium chloride with a concentrated solution of potassium or sodium nitrite as recommended by Corenwinder (1849). Nitrogen is also made by heating ammonium dichromate: $(\text{NH}_4)_2\text{Cr}_2\text{O}_7 = \text{Cr}_2\text{O}_3 + 4\text{H}_2\text{O} + \text{N}_2$; or a mixture of ammonium chloride and potassium dichromate; or a mixture of hypobromites or hypochlorites with urea; etc. Atmospheric nitrogen is made on a large scale by the fractional rectification of liquid air as indicated on p. 153, and it is used in the manufacture of calcium cyanamide, *q.v.*

§ 2. The Properties of Nitrogen.

What of nitrogen? Is not its apparent great simplicity of action all a sham?—M. FARADAY.

Nitrogen is an odourless colourless gas, not quite so heavy as air. It is slightly soluble in water; 100 volumes of water at 0° absorb 2.4 volumes, and at 20°, 1.6 volumes of gas. At 3500° it is estimated that about 5 per cent. of the nitrogen is dissociated into atoms: N_2 (95 per cent.) = 2N (5 per cent.). Nitrogen has been condensed to a colourless liquid which boils at about -195° at ordinary atmospheric pressure; and solidifies to a white snow-like mass melting at -214°. Solid nitrogen exists in two forms with a transition temperature -237.5° and the molecular heat of transformation 53.8 Cals. Nitrogen cannot be a poisonous gas, for the air we breathe contains a large proportion. The nitrogen dilutes the oxygen as indicated on p. 926. Animals die in nitrogen owing to suffocation, *i.e.* want of oxygen necessary for respiration. Nitrogen is both incombustible and a non-supporter of ordinary combustion.

Molecular nitrogen chemically inert; atomic nitrogen active.—The chief characteristic of nitrogen gas is its chemical inertness, due, it is sometimes stated, to "the great affinity of the atoms in the molecule for one another." At any rate, when nitrogen is combined with other elements the converse is true, for the nitrogen compounds generally possess great chemical activity. Witness nitrogen chloride, possibly the most violent explosive known; potassium nitrate in gunpowder; the white and smokeless powders; and explosives generally—most of them contain nitrogen— NO_2 , or NO , *e.g.* nitroglycerol, picric acid, etc.

Nitrides.—While the idea that nitrogen is an inert gas with a disinclination to enter into chemical action is now said to be "an ancient fiction to be relegated to the archives of science," it is still true that nitrogen is inert in the sense that it is difficult to initiate reactions in which nitrogen takes part, and it is usually necessary to raise the system to a

high temperature to inaugurate the reaction, showing that nitrogen leaves its state of "chemical rest" with difficulty. Nitrogen combines directly with a few elements under special conditions, *e.g.* with calcium, beryllium, magnesium, lithium, titanium, etc., when these metals are heated in an atmosphere of nitrogen. Aluminium, for instance, absorbs 12.2 per cent. at 900°. If a mixture of alumina and carbon is heated to about 1800° in a current of nitrogen, aluminium nitride is formed while the carbon and oxygen produce carbon monoxide: $\text{Al}_2\text{O}_3 + 3\text{C} + \text{N}_2 = 2\text{AlN} + 3\text{CO} - 187.6 \text{ Cals.}$ The following experiment illustrates the reaction:

Heat a mixture of aluminium powder and 2 or 3 per cent. of lamp black in a closed crucible. When the crucible is white-hot, lift up the lid. A mere surface crust reacts with the oxygen of the air, while the nitrogen penetrates downwards into the interior of the crucible forming aluminium nitride. When the crucible is cold, fine bluish crystals of the nitride will be found inside.

The compounds of the metals with nitrogen are called nitrides, just as the corresponding compounds with oxygen are called "oxides." Hence, when some of these metals are oxidized in air, a little nitride as well as oxide may be formed. Most nitrides furnish ammonia when treated with water, so that the smell of ammonia may sometimes be detected when the skimmings or dross from molten metals are treated cold, with water. The formation of magnesium nitride can be illustrated by passing nitrogen through the apparatus illustrated in Fig. 226 in the direction *A.B.* A quartz or hard glass tube contains a boat with a little powdered magnesium. This tube is connected with a glass bulbed tube dipping in coloured water to serve as gauge. When all the air has been expelled by the nitrogen, close the stopcock *A*, heat the tube with a Bunsen's burner, and finally with a large Meker burner, or a blast gas burner. When the temperature reaches 900° the bubbling of gas from the tube by thermal expansion will cease, and the rise of the coloured water in the gauge *B* indicates that the metal is absorbing the gas. The nitrides of many elements are formed by heating them in ammonia gas. Iron nitride, Fe_3N_2 , is obtained by heating finely divided iron in a rapid current of ammonia between 450° and 475°; it is completely decomposed at 600° to 700°. When the nitride is heated in nitrogen alone below 600° it loses about half its nitrogen, probably forming a more stable nitride, Fe_4N . With acids the nitride forms ferrous salts, ammonium salts, and hydrogen. In view of these properties, G. Charpy and S. Bonnerot conclude that the formula

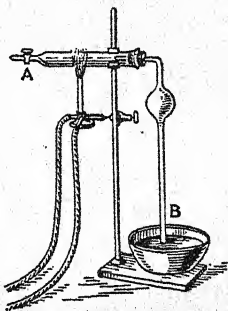
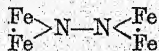


FIG. 226.—Formation of Nitrides.



best represents the constitution of the compound. The nitrogen found in metallic iron is considered to be occluded, or else combined with some element other than iron. The heats of formation of the nitrides are comparatively large, thus, for Mg_3N_2 , the heat of formation is 119.7 Cals., for Ca_3N_2 , 111.2 Cals., and for AlN , 55 Cals.

Allotropic nitrogen.—About 1820, there was much discussion as to

whether nitrogen is an element or a compound. Berzelius, for instance, considered nitrogen to be a suboxide of an unknown element which he called "nitricum," but that view was not consistent with the definition of an element. So also allotropic modifications of nitrogen have been reported from time to time, but the allegations have not been established. In 1910 R. J. Strutt noticed that "vacuum tubes" containing nitrogen—freed from oxygen by previous exposure to phosphorus and red-hot copper—when subjected to the high-tension electric discharge from a Leyden jar, show an "after-glow" when the discharge is stopped. The experiment can be illustrated by Fig. 227. The tube *AB* is supposed to contain nitrogen. The discharge is passed in the direction indicated by the dotted lines shown between the terminals, while the nitrogen travels into *B*, where it appears as a "whirling cloud of brilliant yellow light." The nitrogen is "activated" by the discharge; for the gas then appears to be in a more active condition chemically than ordinary nitrogen. The activity is not due to the presence of ions produced by the discharge, because the chemical activity of the nitrogen persists after the ions are removed. The presence of oxygen or water vapour seems to retard the formation of the allotropic nitrogen. The "after-glow" which accompanies the conversion of the chemically active modification into ordinary nitrogen is intensified

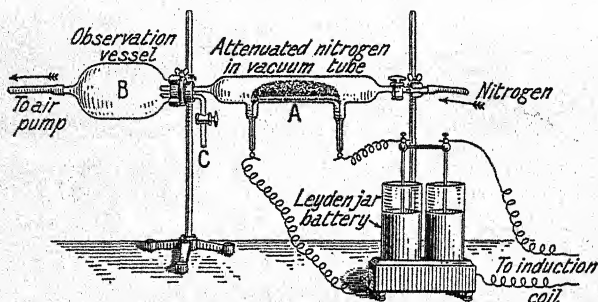


FIG. 227.—Formation of Strutt's Allotropic Nitrogen.

by cooling and weakened by heating. The presence of oxygen destroys the activity, hydrogen has no action. The active nitrogen gradually returns to normal nitrogen on standing, and very readily in the presence of oxidized copper. The active modification combines with phosphorus at the same time the excess of phosphorus is converted into the red variety;¹ the "active" nitrogen also combines with sodium and mercury, forming nitrides. These are decomposed by water or an aqueous solution of potassium hydroxide forming ammonia. The mercury nitride is explosive. Stannic and titanous chlorides furnish white deposits containing nitrogen; and with nitric oxide it forms nitrogen peroxide—apparently by the very curious reaction: $2\text{NO} + \text{N} = \text{NO}_2 + \text{N}_2$. A greenish-yellow flame is developed during the last-named reaction; the flame resembles that produced when ozone reacts with nitric oxide. The tube *C*, Fig. 227, enables gases and vapours to be brought into contact with the active nitrogen in *B*.

¹ It is not uncommon to find some phosphorus converted into red phosphorus when a mass of phosphorus reacts with another substance.

When an acetylene or chloroform flame is fed with a little active nitrogen, the spectrum in each case is characteristic of cyanogen; solid and liquid products are obtained, as well as hydrogen cyanide (no cyanogen): $C_2H_2 + 2N = 2HCy$. Similar results were obtained in the combustion of other organic vapours. If halides are present—e.g. carbon tetrachloride—then cyanogen chloride is formed; with benzene, cyanobenzene is produced. When mixed with the vapour of carbon disulphide, nitrogen sulphide, $(NS)_n$, and brown carbon monosulphide are produced: $CS_2 + N = NS + CS$. Hydrogen sulphide and sulphur chloride also form nitrogen sulphide. Strutt, the discoverer of these phenomena, suggested as a trial hypothesis, that the "active modification of nitrogen" is nitrogen in the atomic condition, but it cannot yet be said that the action is clearly understood.

Atomic weight of nitrogen.—The "combining weight" can be determined by finding the amount of hydrogen or oxygen which combines with a known weight of the gas. For example, the analysis of nitrous oxide, nitric oxide, ammonia, etc., furnishes: $O:N = 16:14\cdot01$. Collect together the vapour densities of all the known volatile compounds of nitrogen. We obtain a table from which Table XXXVI. has been abridged.

TABLE XXXVI.—VAPOUR DENSITIES OF VOLATILE NITROGEN COMPOUNDS.

Volatile compound.	Vapour density.	Formula of compound. Molecular weight = vapour density.	Amount of nitrogen in the molecule.
Ammonia	17·03	NH_3	14·01
Hydrazoic acid	43·04	N_3H	42·03
Nitric oxide	30·01	NO	14·01
Nitric peroxide	46·01	NO_2	14·01
Nitrogen	28·02	N_2	28·02
Nitrous oxide	44·02	N_2O	28·02

The smallest amount of nitrogen entering into the composition of any of these molecules lies somewhere between 14·00 and 14·15; the best representative value is taken to be 14·01, when the atomic weight of hydrogen is 1·008, and of oxygen 16. Hence this number represents the atomic weight of nitrogen.

§ 3. The Composition of the Atmosphere.

I conceive the confused mass of vapour, air, and exhalations which we call the atmosphere, to be nothing else but the particles of all sorts of bodies of which the earth consists, separated from one another and kept at a distance by repulsive forces.—ISAAC NEWTON (1717).

The air contributes to so great a number of chemical phenomena by the combinations which it forms, that it is important to have a precise idea of the parts which compose it, and of the proportions in which its elements are found.—C. L. BERTHOLLET (1803).

About the middle of the 17th century, Robert Boyle said: "The air is a confused aggregate of effluvioms from such differing bodies, that, though they all agree in constituting by their minuteness and various

motions one great mass of matter, yet perhaps there is scarcely a more heterogeneous body in the world," and these words forcibly emphasize the fact that air is a mixture of several different gases—oxygen and nitrogen along with much smaller quantities of ammonia and other nitrogen compounds: hydrogen, hydrocarbons, hydrogen peroxide, carbon dioxide, sulphur compounds, organic matter, suspended solids, chlorides, ozone, water vapour, argon, helium, krypton, neon, xenon, etc. The more carefully the composition of air is examined, the more do we realize how apposite are Boyle's remarks. The five gases last-named are sometimes called the "noble gases" or the "inert gases" of the atmosphere, and they are generally included with the "atmospheric nitrogen." Dust is also common in air. Thus the outside air in London contains from 80,000 to 116,000 particles per c.c., while inside a room as many as two millions may be present. Over the Pacific ocean from 280 to 2200 per c.c. have been counted. Micro-organisms of various kinds abound in the lower strata of the atmosphere. These cause putrefaction, fermentation, and pathological phenomena. When filtered through asbestos or cotton-wool, air is freed from dust particles, and when a beam of light is passed through dust-free air, it will no longer reveal a multitude of motes constantly in motion. The space is optically empty.

Oxygen and nitrogen.—Some of the early workers believed that the proportion of oxygen and nitrogen in air varied with respect to time and place, as well as the state of the weather; but the analyses of Cavendish, Macarty, S. H. Davy, and C. L. Berthollet rendered it highly probable that only a minute difference is to be found in the proportions of these two elements with air collected in different localities, and this is confirmed by analyses conducted more rigorously with all the refinements known to chemists. It appears, as C. L. Berthollet recognized in 1803, that the variations reported by the early chemists were solely due to the uncertain action of the agents employed for removing the oxygen from the nitrogen. The following analyses are quoted to illustrate the percentage amount of oxygen in air:—

Locality.	Minimum.	Maximum.	Mean.	Number of analyses.	Analyst.
Scotch Hills.	20.80	21.18	20.97	34	R. A. Smith
Paris	20.913	20.999	20.96	100	V. Regnault
Dresden . . .	20.88	20.97	20.93	46	W. Hempel
Cleveland, Ohio	20.90	20.95	20.93	45	E. W. Morley

Hence, after making due allowance for differences in the methods of analysis by different men, it is clear that the relative proportions of nitrogen and oxygen in the air collected near sea-level are almost, but not quite, constant. Air at higher altitudes has probably quite a different composition. It has been observed that the diffusion in the higher strata is not the same as in the lower strata, and that the heavier gases tend to concentrate near the earth's surface, while the lighter gases accumulate higher up; but there is not sufficient material to draw any further conclusions. The observed intensification of the hydrogen line would not justify the assumption of a hydrogen atmosphere. A. Weneger (1912) estimated that—

Altitude kilometers.	Atmospheric pressure.	Hydrogen.	Helium.	Nitrogen.	Oxygen.	Argon.
0	760	0.0033	0.0005	78.1	20.9	0.937
100	0.0128	67	4	1.0	0	—
500	0.0016	7	—	—	—	—

or graphically as in Fig. 228. He assumed that in the outermost layers a new gas geo-coronium predominates, as is evidenced by the spectra of aurora and of the corona of the sun during an eclipse.

It is sometimes stated that the nitrogen of the atmosphere merely dilutes the oxygen, and so modifies its action on living organisms, and on combustion; that nitrogen is as water to the wine of oxygen, serving to

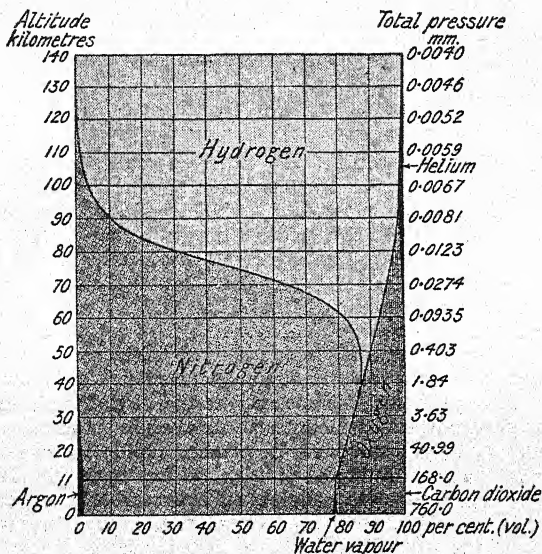


FIG. 228.—Graphic Illustration of the Composition of the Atmosphere at Different Altitudes.

abate the fiery energy of its companion element. It must be remembered that the quantity of oxygen in a given space would not be materially altered if all the nitrogen could be suddenly removed. The particles of one gas do not compress those of another. The distribution of the oxygen is quite independent of that of the nitrogen (Dalton's law p. 101).

Carbon dioxide.—Similar remarks apply to the amount of carbon dioxide. This is rather higher in towns than in the open country; but diffusion of air by winds, etc., prevents an excessive accumulation in any part—excluding, of course, badly ventilated rooms. Thus, J. Reiset (1882) found 3.027 volumes of carbon dioxide per 10,000 volumes of air in Paris; and near Dieppe, 2.942 volumes. These numbers may be regarded as normal. The air over the Antarctic ocean contains about 2.05 parts of carbon dioxide per 10,000 parts of air. This is rather less

than has been noted elsewhere. The proportion of carbon dioxide in air at high altitudes is less than near sea-level. In towns, during a fog, seven or eight volumes may accumulate; and in badly ventilated rooms, ten times the normal amount of carbon dioxide may be present. The other constituents—excluding moisture—are usually regarded as impurities. The essential constituents of normal or average air occur in the following proportions:—

TABLE XXXVII.—AVERAGE COMPOSITION OF ATMOSPHERIC AIR.

Per cent. of	By weight.	By volume.
Nitrogen	75.51	78.03
Oxygen	23.15	20.99
Inert gases	1.30	0.95
Carbon dioxide	0.04	0.03

Ozone, hydrogen peroxide, and nitrogen oxides.—The ozone and hydrogen peroxide are probably formed by electrical discharges in the atmosphere as indicated previously. The same remark applies to the oxides of nitrogen. Free nitric acid has been reported in the atmosphere of tropical regions, but generally, the nitric acid is combined with ammonia. According to A. Levy (1889), about 3 lbs. of ammoniacal nitrogen, and 1 lb. of nitric acid is returned to the earth per acre per annum with the rain. In rural districts the soil is said to receive between 4 and 6 lbs. of combined nitrogen per acre per annum from the rain.

Ammonia.—The ammonia in the atmosphere is largely a product of organic decomposition, and it is returned to the earth by rain in the form of ammonium nitrate, and sometimes as ammonium sulphate or chloride.

Hydrogen and hydrocarbons.—A. Gautier (1901) found that the air of Paris contains per 100 litres—19.4 c.c. of free hydrogen, 12.1 c.c. of methane, 1.7 c.c. of benzene and related hydrocarbons, and 0.2 carbon monoxide with other hydrocarbons. Gautier's estimate is probably rather high. H. Henriet (1904) found 2 to 6 grams of formaldehyde per 100 cubic metres of air. The presence of hydrocarbons explains the oleaginous character of the deposits which form on roofs, leaves of trees, etc., in towns.

Sulphur compounds.—A. Ladureau (1883) reported 1.8 c.c. of sulphur dioxide per cubic metre in the air of Lille. Sulphur compounds are present in small quantities as hydrogen sulphide, sulphur dioxide, and sulphuric acid in the air of towns. According to R. Warrington (1887), about 17½ lbs. of sulphur trioxide is annually "poured" upon each acre of land at Rothamsted. G. H. Bailey (1892) reported a maximum of 0.0267 gram of sulphur expressed as sulphur trioxide per 100 cu. ft. of air near the surface of the ground in Manchester.

Chlorine compounds.—Rain near the sea brings a certain amount of chlorine derived from the sea water. The proportion of salt in the air is greatest near the sea, and diminishes rapidly farther away from the coast. E. Kinch (1900) found, as an average of twenty-six years' observations at Cirencester, that 36.1 lbs. of sodium chloride per acre were brought to the earth with the rain. The amount of "wind-borne" sea salt is greatest

when the wind blows from the sea. Free hydrochloric acid derived from manufacturing operations is sometimes found in the air of towns.

Moisture.—The average amount of moisture, aqueous vapour, in air is rather less than one per cent. by volume; it may reach 4 per cent. in humid climes. The actual amount of aqueous vapour air can carry before it is saturated depends upon the temperature. The higher the temperature, the greater the amount of moisture air can carry (p. 188). Air seldom contains less than 75 per cent. of the possible amount it is capable of holding; but much depends on the local conditions—*e.g.* in the Libyan desert it contains but 9 per cent. of the possible amount. The methods for the determination of moisture in air—hygrometry—are discussed in textbooks on physics.

§ 4. Is Air a Mixture or a Compound of Oxygen and Nitrogen?

Let us assume the function of a judge in a law court, and sum up the evidence for the jury.

1. The proportions of the constituents of air vary a little in different localities, but even this small variation is not found with pure chemical compounds—law of constant proportions, pp. 27, 30. Hence not all the nitrogen and oxygen are combined.

2. The atomic proportion of nitrogen and oxygen in air is as 3·67 : 1; this is approximately as 15 : 4. Hence if all the nitrogen and oxygen are combined, the formula of the compound is $N_{15}O_4$, or $NO_{0.26}$, which does not fit very well with the facts summarized by the law of multiple proportions, p. 40. A similar result is obtained by considering the volume relations of nitrogen and oxygen in air—Gay-Lussac's law, p. 72.

3. The characteristic physical properties (refractive index, absorption of radiant heat, etc.) of nitrogen and oxygen are modified in air only so far as obtains when nitrogen and oxygen are *mixed* in the same proportions. The properties of the two gases are not changed so much as would be expected if a chemical compound were formed. Thus, (a) Biot (1806) found that the refractive index of air is equivalent to that calculated for an equivalent mixture of oxygen and nitrogen; (b) J. Tyndall (1861) observed that the amount of heat absorbed when rays from a given source are allowed to pass through a column of an elementary gas is usually much less than is the case with compounds. Thus air, oxygen, and nitrogen have the same absorptive power, whereas carbon monoxide has an absorptive power of 972 units; nitric oxide, 1590 units; nitrous oxide, 1860 units; and ammonia 5640 units.

4. No heat, no change of volume, or any other sign of chemical change is observed when air is made artificially by mixing the gases together in the right proportions. "If a measurable physical property were different in air and in an equivalent mixture of the constituents of air, the conclusion would follow that air is a compound" (H. St. C. Deville, 1864).

5. The constituents of air can be separated by mechanical means: *e.g.* solution in water (p. 828); by atmolysis (p. 129); by allowing liquid air to vaporize, when the nitrogen distils off before the oxygen (p. 152); and J. Dewar has shown that when solidified air is exposed to a magnetic field, the oxygen is sucked out of it towards the magnet's poles "so that

solid air appears to be a magnia of nitrogen and oxygen "separable by a magnet.

Not one of these five reasons is in itself conclusive, but all, taken together, form a long chain of circumstantial evidence which would lead an unbiased jury to return the verdict: Air is a mechanical mixture of nitrogen, oxygen, etc.

§ 5. The Analysis of Air.

In order to obtain a thorough knowledge of a great number of phenomena it is important to ascertain, with all the precision that can be attained, what are the proportions of oxygen and nitrogen which enter into the composition of the atmosphere.—C. L. BERTHOLLET.

Gravimetric process.—The gravimetric analysis of air was made by J. B. A. Dumas and J. Boussingault (1841) in an apparatus similar in principle to that illustrated in Fig. 229 where a modern furnace is shown in place of the old charcoal furnace, and the number of drying tubes has been reduced. A large globe was evacuated, closed, and weighed in that condition. This globe was connected as indicated in Fig. 229, with a tube

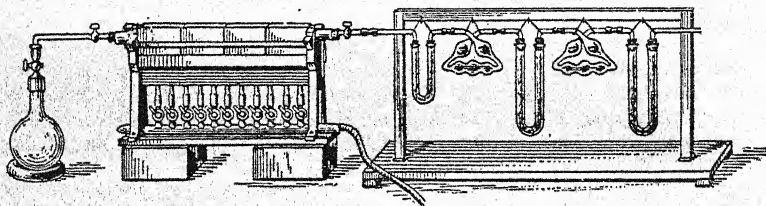


FIG. 229.—Diagram illustrating the Principle of Dumas and Boussingault's Method for the Gravimetric Analysis of Air.

containing metallic copper also evacuated, closed, and weighed. The copper tube was connected with a series of bulbs, and tubes containing concentrated sulphuric acid to remove moisture and ammonia from the air; and with potassium hydroxide to remove carbon dioxide—only a few of the tubes used by Dumas and Boussingault are shown in the diagram. There may be a tainin tube to show that the air which passes into the hot tube is free from carbon dioxide and moisture. The tube containing the copper was heated red hot, and air allowed to enter slowly by gradually opening the stopcocks. The air on its way to the glass globe was deprived of all but the nitrogen (atmospheric). When the globe was full, the apparatus was cooled, and the globe and copper tube weighed. The increase in weight gave the amount of nitrogen in the globe. The tube was also weighed. Its increase in weight represented the weight of oxygen which was associated, as air, with the nitrogen in the glass globe; the tube also contained some nitrogen. The tube was therefore exhausted and weighed again. The difference between the second and third weighings of the tube was taken to represent the nitrogen which must be added to the nitrogen in the globe. The amount of oxygen was calculated from the difference between the first and last weighings of the tube, thus, Dumas and Boussingault obtained the following results:

	Grams.
Tube after experiment	651.415
Evacuated tube before experiment	647.666
Evacuated tube after experiment	651.346
Nitrogen	0.069
Total oxygen in tube	3.680
Globe with nitrogen at 19°, 762.7 mm.	1403.838
Evacuated globe at 19.4°, 762.7 mm.	1391.534
Nitrogen in balloon	12.304
Nitrogen in tube	0.069
Total nitrogen	12.373

Hence in air, the oxygen is to the nitrogen as 3.682 : 12.373, that is as 22.92 : 77.08. As a mean of six determinations they obtained 23.005 grams of oxygen per 76.995 grams of nitrogen. The gravimetric process is very exact. The error need not exceed 0.00001th part of the whole; but the experiment requires special apparatus, and occupies much time. Volumetric processes are not so exact, but they occupy far less time. The methods which have been employed for measuring the relative proportions of oxygen and nitrogen in air are mainly gravimetric and volumetric. Historically, the latter methods were developed before the former. In the volumetric processes, the oxygen is usually removed from the nitrogen by absorbents—alkaline sulphides, phosphorus, alkaline pyrogallol, hot copper, nitric oxide, and water, etc.—which do not disengage any gas, and have no action on the residual nitrogen.

Volumetric processes.—The methods employed by Lavoisier, indicated on pp. 22 and 23, are only rough approximations. J. Priestley and C. W. Scheele simultaneously observed the quantitative composition of air, but the latter first attempted a systematic study of its composition over extended periods of time. He absorbed the oxygen by exposing a confined volume of air to a mixture of iron-filings and potassium sulphide. Priestley's nitric oxide method was much used in the early days.

Cavendish, for instance, is said to have made over 500 analyses by his method before 1790. Hempel's method, with an absorption pipette, charged with sticks of phosphorus, Fig. 230, gives better results. C. W. Scheele used the phosphorus process in 1779. If the pipette be charged with a solution of

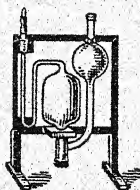


FIG. 230.—Absorption Pipette.

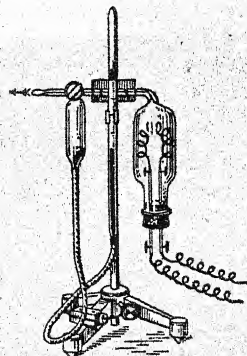


FIG. 231.—Jolly's Apparatus for the Volumetric Analysis of Air.

pyrogallol in one of alkali hydroxide, the oxygen is absorbed and the solution becomes dark brown. This method was used by J. von Liebig in 1851. After the air has been measured in Hempel's burette, Fig. 34, it is transferred to the pipette where the oxygen is absorbed, and then transferred back to the burette for remeasurement. The explosion process indicated on p. 634, for nitrous oxide, may also be employed. The

explosion process (p. 69) first employed by A. Volta in 1774 was tested by Cavendish who obtained very accurate results.

An apparatus, Fig. 231, similar to that designed by P. von Jolly (1879) can be used. A glass globe is provided with an arrangement whereby a piece of copper wire can be heated red hot while inside. The globe is fitted with a three-way stopcock which connects it with the gas-measuring tubes or with the outside air. The gas-measuring tubes are filled with mercury, and the globe is then connected with the air-pump, and exhausted. A jacket containing iced water is placed around the globe. Air dried by passage through wash-bottles containing sulphuric acid is allowed to enter the globe, which is then put in communication with the gas-measuring tube, and the level of the mercury in the levelling tube is read when the mercury in the other tube reaches the mark just below the stopcock. Call this position p_0 . The cold jacket is then removed, and the copper wire heated. When the oxygen is all converted into copper oxide, again place the cold jacket in position, and alter the levelling tube until the mercury reaches its former position. The height of the mercury in the levelling tube is again read. Call this position p . The pressure of the mercury has been reduced $p_0 - p$ owing to the removal of oxygen. Consequently, since the temperature is the same in both cases, the percentage x of oxygen in the air must be $p_0 : p_0 - p = 100 : x$.

§ 6. Argon and its Companions.

Atomic weight, $A = 39.9$; molecular weight, $A = 39.9$. Melting point, -187.9° ; boiling point, -186.1° ; critical temperature, -117.4° . Vapour density ($O_2 = 32$), 39.914; (air = 1), 1.3785. A litre of argon under normal conditions weighs 1.78376 grams.

The discovery of argon.—During 1893–95 Rayleigh found that the density of nitrogen derived from the atmosphere by removing the impurities, ammonia, carbon dioxide, moisture, and also the oxygen, was 14.070 (hydrogen unity); and when the nitrogen was prepared from nitrous oxide, nitric oxide, ammonium nitrite, urea, or magnesium nitride, the density was 14.005. The difference is much larger than the experimental error involved in the determination, and it was therefore inferred that the nitrogen of the air must contain another gas—allotropic nitrogen, or some new gas—previously overlooked. The disturbing gas was isolated by W. Ramsay in May, 1894. It proved to be a new gas which was named argon—from the Greek *ἀργος* (argos), inert, idle.

Isolation of "argon."—The new gas can be obtained by passing atmospheric nitrogen over heated magnesium, or over a mixture containing one part by weight of magnesium filings, 0.25 sodium, and five of quicklime. The magnesium reacts with the quicklime, liberating calcium, which absorbs the nitrogen, and also any oxygen which may be present. Argon can also be prepared by passing air through a hot tube containing calcium carbide at about 800° . The oxygen is absorbed: $2CaC_2 + O_2 = 2CaO + 4C$; and the nitrogen forms cyanamide: $CaC_2 + N_2 = CaCN_2 + C$. The carbide is said to be much more active if it be mixed with 11 per cent. of calcium chloride.

Rayleigh also passed a series of electric sparks through a mixture of air and oxygen (p. 617), and removed the nitrogen oxides as fast as they

were formed by alkalis. About $\frac{1}{4}$ of the atmospheric nitrogen remains as a residue which will not combine with the oxygen when sparked in the presence of alkalis. This process for isolating the inert gas from air is particularly interesting because H. Cavendish made a similar experiment in 1785. He laboriously eliminated all the known constituents of air, but the air did not all disappear; a remnant "stood its ground." He concluded that if any part of the nitrogen of atmospheric air differs from the rest "we may safely conclude that it is not more than $\frac{1}{120}$ of the whole." Cavendish did not pursue the matter further; he was satisfied to find it so small. Remembering the conditions under which Cavendish's experiment was made, this estimate is remarkably close to Ramsay's number, $\frac{1}{41}$. Cavendish's observation was overlooked until Rayleigh and Ramsay indicated its importance.

Properties of argon.—Argon gas is without colour, taste, or smell. It is incombustible and a non-supporter of combustion. Argon is remarkable in forming no well-defined compounds with other elements. There are reports of combinations with benzene vapour and with carbon disulphide, but these have not been substantiated. So far as we know, its chemical indifference is absolute; it is outside the pale of orthodox chemistry. 100 volumes of water at 0° dissolve 4.47 volumes, and at 20°, 2.86 volumes. Hence argon is rather more soluble than air in water, and in consequence, when the gases dissolved in rain water are expelled by boiling, the resulting "air" is slightly richer in argon than ordinary air. Argon is present in the gases from sea water, rivers, and springs, as well as occluded in certain minerals—clèveite and uraniferous minerals—and in some meteorites. It also occurs in rock salt springs, where it must have been immersed for ages. The gas must have been captured from the diluvian atmosphere. The argon which was found occluded in meteorites could not be induced to re-enter the aerolite. The conditions under which the gas was imprisoned cannot be reproduced. Although the atmosphere contains so small a percentage—less than one per cent.—the absolute quantity is enormous—every square mile of terrestrial surface is loaded with about 800,000,000 lbs. of argon.

The inert gases.—Argon isolated from atmospheric air as described above is contaminated with minute quantities of some other gases, for if the liquefied gas be fractionally distilled, Ramsay found that four other gases could be obtained: helium—named from Greek *ἥλιος* (helios), sun; neon—from *νέος* (neos), new; krypton—from *κρυπτός* (kryptos), hidden; and xenon—from *ξένος* (xenos), stranger. The neon, helium, argon, and contaminating nitrogen pass off first; the xenon and krypton remain. The more volatile bases are compressed in a vessel cooled by liquid air. By repeated fractionation of the less volatile residual liquid, the xenon and krypton can be separated from the argon and from one another. Similarly, neon and helium can be separated from the contaminating nitrogen and argon. By surrounding the mixture of helium and neon with liquid hydrogen, the neon freezes to a white solid from which the helium can be removed by the air pump. The physical properties of these gases are summarized in Table XXXVIII. (p. 690).

Like argon, they are all inert chemically, no well-defined compounds with other elements have been obtained. Hence it is generally stated that these elements are null-valent. These gases can be excited elec-

trically so as to furnish characteristic and complex spectra. Ordinary molecular nitrogen, as we have seen, is somewhat inert to ordinary chemical influences, while in the atomic condition it is one of the most active of elements. Accordingly, H. E. Armstrong (1895-1912) argued that it is possible that the molecules of the inert gases are made up of two or more atoms so firmly bound together that they cannot be severed by any known chemical process. The evidence for the familiar statement: "The molecules of the inert gases are monatomic," as we shall soon see, is not conclusive. Neon has been shown, by atmolysis, to be a mixture of two other gases of atomic weight 19.9 and 22.1 respectively, but no difference has been detected in the physical properties of the two fractions, *vide* isotopes. There is no perceptible difference in the spectra of the two components of neon. The inert gases can be obtained in relatively large quantities as a by-product in the manufacture of liquid air. It has been proposed to use neon along with a little mercury vapour in tubes for illuminating purposes, because (1) a comparatively small amount of electrical energy is necessary to make the tube luminous—it is said that 1000 volts are needed with air under conditions where but 13 volts are required for neon. It is also claimed (2) that in the light emitted from the neon lamp blue, red, and green have apparently the same colours as they have in daylight.

TABLE XXXVIII.—PHYSICAL PROPERTIES OF THE NULL-VALENT OR INERT GASES.

Gas.	1000 vols. air con- tain approx- imately.	Density $O_2 = 32$ atomic [and molecular] weight.	Melting point.	Boiling point.	Critical tempera- ture.
Helium, He. . . .	0.0014	3.99	—	-268.7°	-268°
Neon, Ne	0.015	20.20	-253°	-233°	-220°
Argon, A	9.37	39.88	-188°	-186°	-117°
Krypton, Kr . . .	0.00005	82.92	-169°	-152°	- 63°
Xenon, X	0.000006	130.22	-140°	-109°	+ 15°

Helium.—In 1868 P. J. C. Janssen detected a prominent orange line in the spectrum of the sun's photosphere. This did not correspond with the spectral lines of any known element. Hence E. Frankland and J. N. Lockyer postulated the existence of a new element which they called helium. The same orange line was later detected in the spectrum of certain stars; and in 1882 Palmieri noticed the same line in the spectrum of the gases flaming from Vesuvius. While seeking for occluded argon in the "nitrogen" which had been reported in many minerals—*clèveite*, *fergussonite*, *bröggerite*, *uraninite*, etc.—W. Ramsay found that the gas was neither argon nor nitrogen. The spectrum of the new gas was identical with that reported by Janssen and by Palmieri. Hence its name, helium. The same gas was later detected in atmospheric air as indicated above; in the gases dissolved in certain mineral waters; and also in the gaseous emanations from certain springs—*e.g.* the gas from the Sautenay springs

has 8-10 per cent. of helium; that from Maizières, 6 per cent.; that from Grisy, 2.2 per cent.; and that from La Bourboule, 0.1 per cent. As in the case of argon, when the helium imprisoned in minerals is once liberated it cannot be re-incarcerated. The gas seems to be generated in these minerals as a product of the radioactive decay of some of the constituents of the mineral—*vide* radioactivity. By the rapid evaporation of liquid helium a temperature within 2° of absolute zero -273° has been obtained by H. K. Onnes (1908). For the use of helium in aeronautics, *vide* hydrogen.

Removal of the gases occluded in minerals.—To remove the occluded gases from minerals, the pulverized mineral is heated in a hard glass tube, *A*, Fig. 232, with or without concentrated sulphuric acid, or sodium pyrosulphate. The system is fitted as shown in Fig. 232, and exhausted by means of an air pump; Toepler's mercury pump,¹ *C*, is shown in Fig. 232. The mineral is heated under reduced pressure, and the gases are collected in *D*, the receiver of the pump. The gases on their way to the pump travel through a flask, *B*, containing sticks of solid potassium hydroxide to absorb water and sulphur compounds. The gas collected in the gasholder *D* can then be analyzed, or freed from hydrogen and hydrocarbons by passage over hot copper oxide; from oxygen by passage over hot copper; from carbon dioxide by soda lime, or potassium hydroxide; from nitrogen by hot magnesium; and from moisture by phosphorus pentoxide—and the residue examined by the spectroscope for helium, etc. The helium occluded in monazite sand can be readily obtained by this process.

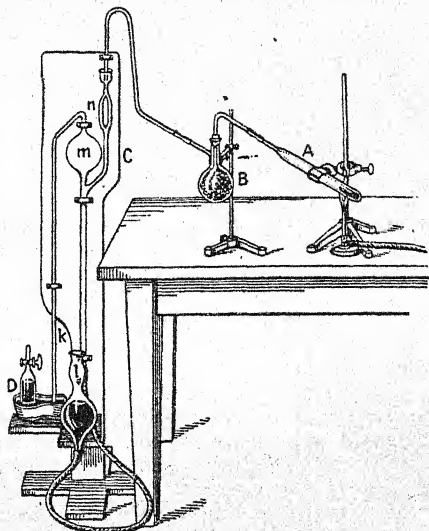


FIG. 232.—Removal of Occluded Gases from Minerals.

§ 7. The Two Specific Heats of Gases.

It will be remembered that "specific heat" is a term employed to represent the amount of heat required to raise the temperature of one

¹ Toepler's pump, *C*, is worked by raising the levelling tube, *l*; air is expelled from the exit tube, *k*, by the mercury rising in *m*. The levelling tube is again depressed, part of the air in *AB* enters the globe *m*. The air in *m* is expelled by again raising the levelling tube *l*, until a little mercury runs down the tube *k*; the valve *n* prevents running back into *B*. These operations are repeated until the required degree of exhaustion has been attained.

gram of a substance 1° . A gas can be heated by simple compression, its specific heat must then be zero; but a certain amount of energy, equivalent to the specific heat, is needed for the work of compression. Again, a gas, if it be expanded, is cooled; if the cooling effect of expansion just counterbalances the heat added to the gas, the temperature remains constant; and the specific heat appears to be indefinitely large. Here work, equivalent to the heat supplied, is performed by the expanding gas. These facts show that the condition of the gas must be stated before it is possible to define its specific heat. It is conventionally agreed that if the gas be allowed to expand during a change of temperature so that its pressure remains constant, the amount of heat required to raise the temperature of one gram of the gas 1° shall be called the specific heat under constant pressure, and symbolized by C_p . If the pressure be increased so that the volume remains constant when the gas is heated, the amount of heat required to raise the temperature 1° is likewise called the specific heat under constant volume, and symbolized C_v .

In the following discussion, it will be remembered that the kinetic theory assumes that the temperature is proportional to the average speed of translation of the moving molecules—an increase of the speed is accompanied by a rise of temperature, and conversely. The heat imparted to a gas is not spent merely in raising the temperature of the gas; that is, in speeding up the motions of the molecules. Energy is spent in—

(1) **Augmenting the speed of the moving molecules.**—The heat required to actually increase the kinetic energy of the moving molecules so as to produce a rise of temperature is the same for all gases. Let a denote this quantity.

(2) **Performing external work.**—Heat energy is needed to overcome the pressure of the atmosphere when the gas is allowed to expand. Call this quantity b . Since the coefficient of thermal expansion of all gases is the same (p. 103), this quantity is practically constant for equal volumes or equimolecular weights.

(3) **Performing internal work.**—Heat energy is required to produce changes within the molecule which may alter the motions or orientation of the constituent atoms of the molecule, or raise the kinetic energy of the atoms moving within the molecule. Let c denote the energy spent within the molecule per degree rise of temperature. A certain amount of energy must also be spent in overcoming the effects of intermolecular attractions (p. 124). This can be neglected for the time being.

The ratio of the two specific heats may now be written:

$$\frac{C_p}{C_v} = \frac{a + b + c}{a + c}$$

The specific heat of a gas at constant volume.—We have seen, p. 140, that $pv = \frac{1}{2}MV^2$ where M denotes the mass, and V the average velocity of the molecules. But the kinetic energy of a body of mass M moving with a velocity V is $\frac{1}{2}MV^2$; hence $pv = \frac{2}{3} \times \frac{1}{2}MV^2$; or the kinetic energy of the molecular motions is $\frac{3}{2}pv$. But $pv = RT$, p. 105. Hence the kinetic energy of molecular motion is $\frac{3}{2}RT$. If one gram of gas be heated 1° , the kinetic energy becomes $\frac{3}{2}R(T + 1)$. Hence if the gas be heated 1° at constant volume, the thermal value of the increased kinetic

energy is $\frac{3}{2}R(T + 1) - \frac{3}{2}RT = \frac{3}{2}R$ cal. This result represents the specific heat of the gas at constant volume; or, $C_v = \frac{3}{2}R$.

The external work done by an expanding gas.—Again, if a gram of gas expands against atmospheric pressure when its temperature is raised 1° , the gas, in consequence, does work by pressing back the atmosphere, so to speak. The equivalent of this work must be supplied in the form of heat. This work is equivalent to the product of the pressure against the change in volume. Let x denote the change in volume when the gas is heated 1° , under a constant pressure; then, $p(v + x) = R(T + 1)$, and $p v = RT$, p. 105. By subtraction $p x = R$. This means that when a gram of gas is heated 1° , the resulting expansion against atmospheric pressure does work equivalent to R cal.

The specific heat at constant pressure.—Hence, R cal. must be added to the previous result to obtain the thermal equivalent of the energy supplied to one gram of gas in the form of heat when its temperature is raised 1° . Otherwise expressed, if one gram of gas be heated 1° , at constant pressure, an amount of heat equivalent to $\frac{3}{2}R + R = \frac{5}{2}R$ is required. This result represents the specific heat of the gas at constant pressure, or $C_p = \frac{5}{2}R$.

The numerical value of R can be calculated:

From equation (6), p. 105, it follows that $p v / T = p_1 v_1 / T$, and if v_1 represents the volume of a gram-molecule of any gas at N.P.T., we have, p. 84, $v_1 = 22,300$ c.c.; p_1 from p. 95, is 1033.3 grams per square centimetre; and $T_1 = 273^\circ \text{C}$. Hence,

$$R = \frac{p v}{T} = \frac{1033.3 \times 22,300}{273} = 84,760 \text{ gram-centimetres of energy.}$$

From Joule's experiments on the mechanical equivalent of thermal energy it is known that one gram-centimetre of mechanical energy is equivalent to 42,650 calories. Hence, $p v / T = 2$ Cals. nearly; but (5), p. 105, $p v / T = R$; hence, $R = 2$ Cals. nearly.

The ratio of the two specific heats of a gas.—Returning to the ratio of the two specific heats, which is usually symbolized γ , we can now write $a = \frac{3}{2}R$, and $b = R$, or:

$$\gamma = \frac{C_p}{C_v} = \frac{\frac{5}{2}R + c}{\frac{3}{2}R + c}$$

The magnitude of c will vary with different gases, for it will naturally be related somehow with the complexity of the molecule. The greater the value of c , the less the value of the ratio of the two specific heats. For a monoatomic gas, c is probably zero, and the numerical value of the ratio becomes $\gamma = \frac{5}{3}$, or 1.67. The greater the complexity of the molecule, the greater the value of c , and the smaller the value of the ratio of the two specific heats. This is illustrated by the following table:

TABLE XXXIX.—RATIO OF THE TWO SPECIFIC HEATS OF GASES.

Gas.	Mole- cule.	Atoms per mole- cule.	γ	Gas.	Mole- cule.	Atoms per mole- cule.	γ
Mercury . . .	Hg	1	1.67	Carbon dioxide . .	CO ₂	3	1.31
Argon . . .	A	1	1.65	Nitrous oxide . .	N ₂ O	3	1.31
Hydrogen . . .	H ₂	2	1.41	Hydrogen sulphide	H ₂ S	3	1.31
Nitrogen . . .	N ₂	2	1.41	Ammonia . . .	NH ₃	4	1.30
Oxygen . . .	O ₂	2	1.40	Methane . . .	CH ₄	5	1.27
Carbon monoxide	CO	2	1.40	Ethylene . . .	C ₂ H ₄	6	1.24
Hydrogen chloride	HCl	2	1.39	Ethane . . .	C ₂ H ₆	8	1.18
Chlorine . . .	Cl ₂	2	1.32	Alcohol . . .	C ₂ H ₅ OH	9	1.13
Bromine . . .	Br ₂	2	1.29	Benzene . . .	C ₆ H ₆	12	1.09
Iodine . . .	I ₂	2	1.29	Ether . . .	C ₄ H ₁₀ O	15	1.06
Iodine chloride .	ICl	2	1.31	Turpentine . . .	C ₁₀ H ₁₆	26	1.03

The ratio of the two specific heats and molecular weights.—These numbers mean that if the ratio of the two specific heats of a gas be about 1.6, the gas will usually have one atom per molecule, with a ratio about 1.4, two atoms per molecule, and about 1.3, three atoms per molecule. The kinetic theory would have no explanation to offer if the value of γ were greater than $1\frac{2}{3}$; but no cases are known. There are a number of discrepancies. This must be expected owing to differences in molecular attraction, tendencies to polymerization, dissociation, etc., which affect the value of c . The coloured gases—chlorine, bromine, iodine, and iodine chloride, with two atoms per molecule—give lower values than is usually obtained with the colourless diatomic molecules; and gases which are readily condensed to liquids give rather lower values than those less readily liquefied. Hence if the ratio of the two specific heats of a gas falls into one of these groups—1.6, 1.4, 1.3—this fact may be taken as circumstantial evidence, but not conclusive proof, that the molecule has a corresponding number of atoms per molecule. There is, however, no unimpeachable law connecting the specific heat of a complex molecule with the *number* of the constituent atoms which is independent of the *nature* of the atoms. The ratio of the two specific heats of argon and the inert gases appears to be about 1.6, and hence it is supposed that the molecules of these gases are monatomic, like mercury. This means that the density ($H = 2$), the molecular weight, and the atomic weight will probably have the same numerical value. Hence the determination of the ratio of the two specific heats¹ provides an independent method of ascertaining the number of atoms in the molecules of a gas without reference to the compounds of the element. In the case of mercury, the monatomicity of the gas has been established altogether apart from this reasoning.

This subject cannot be passed by without bringing the weak step in the above reasoning into prominence. The low molecular heats of the inert gases are assumed to prove that these gases have one-atom molecules. But it is easy to see that if little or no heat is expended in doing internal work when the temperature of a gas is raised, a gas with poly-

¹ The value of γ can be determined experimentally by measuring the velocity of sound in gases. For this see any text-book on Physics.

atomic molecules might be reported to have monatomic molecules. Unlike mercury, the inert gases do not form chemical compounds, and hence the number of atoms in the molecule cannot be determined by the usual methods based upon Avogadro's hypothesis. The inference that the molecules of the inert gases are monatomic thus involves an assumption which is less readily granted than is the case with mercury, cadmium, etc., which do form volatile chemical compounds.

§ 8. The Specific Heats of Elementary Solids—Dulong and Petit's Rule.

The atomic heat, that is, the thermal capacity of the atoms of an element, is the product of the specific heat and atomic weight of the element. P. L. Dulong and A. T. Petit (1819) in their study of the specific heats of different solid elements obtained a remarkable result. They found: The product of the atomic weight, w , and the specific heat, c , of an element has nearly always the same numerical value—6.4, or say 6—Dulong and Petit's rule. This means that the atomic heats or the thermal capacity of the atoms of the elements are approximately the same. The relation is usually expressed:

$$\text{Atomic heat} = Cw = 6.4$$

In illustration, a few elements may be selected at random from a list containing nearly 50 elements for which data are available:

TABLE XL.—ATOMIC HEAT OF ELEMENTS.

Element.	Specific heat.	Atomic weight.	Atomic heat.
Lithium	0.9408	6.94	6.53
Silver	0.0559	107.88	6.03
Gold	0.0304	197.2	6.25
Copper	0.0923	63.57	5.88
Bismuth	0.0305	208.0	6.34
Lead	0.0315	207.10	6.52
Aluminium	0.2143	27.1	5.81
Iron	0.1098	55.85	6.12
Uranium	0.0277	238.5	6.61

The atomic weights here range from 6.94 to 238.5, and yet, when multiplied by the respective specific heats, the products are nearly constant. Rigorous agreement cannot be expected. The divergencies are too large to be accounted for by the inevitable errors of observation involved in measuring the specific heats, but the very irregularity of the divergencies leads to the view that Dulong and Petit's law approximates to a truth, and that the observed differences are due to disturbing effects which are not functions of the atomic weight—e.g. crystalline form; different relations between the temperature at which the specific heats were determined and the critical fusion temperature; etc.

Influence of temperature on the atomic heats of carbon, silicon, boron, and beryllium.—Silicon, boron, beryllium, and carbon, at ordinary

temperatures, have atomic heats represented respectively by 4.8, 2.6, 3.4, and 1.35; but at higher temperatures, these elements approximate close to 6. This is illustrated by the diagram, Fig. 233, which represents the influence of temperature on the atomic heats of the elements named. The specific heats of silicon, boron, beryllium, and carbon change with temperature, until a point is reached at which they are nearly constant. This point is near 600° for carbon and boron, and about 200° for silicon. Conversely, the specific heats of the metals become very small towards the absolute zero. For example, the amount of heat required to raise the temperature of copper at -250° one degree is thirty times smaller than that required at 0°. The atomic heats of the elements generally, diverge from Dulong and Petit's constant as the temperature falls, and converge towards a constant value as the temperature rises, *e.g.* :

Metal.	-150°	-100°	-50°	0°	50°
Silver . . .	4.97	5.46	5.80	6.03	6.06
Gold . . .	5.25	5.54	5.78	5.97	6.10
Bismuth . .	5.49	5.67	5.86	6.06	6.27
Aluminium .	3.71	4.54	5.19	5.68	5.89

Some believe that the divergencies are mainly due to the fact that the temperatures at which the specific heats have been determined stand in

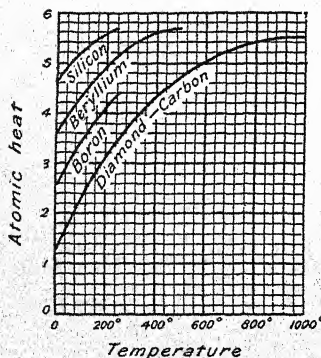


FIG. 233.—Effect of Temperature on Atomic Heat Curves.

raising the kinetic energy of the constituent atoms; in raising the volume of the body; and in overcoming molecular attractions. The coefficient of thermal expansion of solids is small, and therefore also the work of expansion of solids against external pressure is small. The difference between the specific heats of a solid and gaseous element is usually great. Thus iodine, gas, has an atomic heat 3.3; and iodine, solid, 6.9; bromine, gas, 4.7; bromine, solid, 6.7. Variations in the complexity of the molecules of an element lead to similar differences. Thus: the

specific heat of amorphous carbon is 0.2609; graphite, 0.2000; and diamond, 0.1470; the specific heat of ordinary tin is 0.0559, and of allotropic tin, 0.0545. Similarly, differences in the physical condition may also produce an effect on the specific heat. Thus: the specific heat of hard tempered steel is 0.1175, and of soft tempered steel, 0.1165. Consequently, it must be inferred that the heat does important work other than merely raising the kinetic energy of the molecules. It is therefore strange that the relation pointed out by Dulong and Petit does not exhibit greater divergencies.

The rectification of atomic weights by Dulong and Petit's rule: $Cw = 6.4$. It will be obvious that if the specific heat of an element be known, it is possible to compute an approximate value for the atomic weight. The number so obtained may be useful for deciding between two numbers which are multiples of a common factor. The method is obviously only applicable to elements whose specific heat can be determined. W. A. Tilden (1900-3) tried unsuccessfully to find the conditions of temperature for which Dulong and Petit's law holds absolutely. Hence he concluded that the usual application of this law to the rectification of atomic weights "is a rough empirical rule, which, setting aside silicon, boron, beryllium, and carbon, is only available when the specific heats have been determined at comparatively low temperatures, usually and most conveniently between 0° and 100° ." The specific heat method of fixing the atomic weights is not used in modern chemistry because of the variation in the specific heat with temperature, and because more exact methods are available.

EXAMPLES.—(1) What is the atomic weight of silver assuming that the specific heat is 0.0559? Here, $6.4 \div 0.0559 = 108$ nearly. This is close to the accepted value for the atomic weight of this element.

(2) Platinum chloride, on analysis, furnished 35.5 grams of chlorine per 48.6 grams of platinum. The specific heat of platinum is 0.0324, and the atomic weight is approximately $6.4 \div 0.0324 = 197.5$. Hence, since $197.5 \div 48.6 = 4$ (nearly), it follows that if the atomic weight of chlorine is 35.5, the atomic weight of platinum must be nearly $48.6 \times 4 = 194$.

(3) When indium was first discovered the analysis of its chloride furnished indium 37.8; chlorine 35.5. The equivalent of indium is therefore 37.8. The formula of the chloride was thought to be InCl_2 , and the atomic weight was accordingly represented 75.6. The specific heat of the metal was found to be 0.057. Hence, $75.6 \times 0.057 = 4.5$. If 75.6 be the correct atomic weight, the product would approximate closer to 6.4, and hence it was inferred that 75.6 is not the correct atomic weight of indium, rather does the atomic weight approximate to $6.4 \div 0.057 = 112.3$. If InCl_2 be the formula of the chloride, the atomic weight will be $27.8 \times 3 = 113.4$, which is the number usually adopted for the atomic weight of this element.

§ 9. Molecular Heats.

Solids.—The molecular heat or thermal capacity of the molecules of a substance is defined as the product of its specific heat and its molecular weight. In 1831 F. E. Neumann noticed that the product of the specific heat and the molecular weight of compounds of similar composition is nearly constant—Neumann's law. The value of the constant varies from one series of compounds to another. Thus:

TABLE XLI.—MOLECULAR HEATS OF SOLIDS.

Car- bonates.	Mol. wt.	Sp. heat.	Mol. heat.	Chloride.	Mol. wt.	Sp. heat.	Mol. heat.
CaCO ₃	100.09	0.206	20.6	BaCl ₂	208.29	0.090	18.7
SrCO ₃	147.02	0.145	21.3	SrCl ₂	158.54	0.12	19.0
BaCO ₃	197.37	0.109	21.4	PbCl ₂	277.02	0.066	18.3
PbCO ₃	277.02	0.080	21.3	HgCl ₂	270.92	0.069	19.2

J. P. Joule, in 1844, brought forward some evidence indicating that the molecular heat of a solid compound is approximately the sum of the atomic heats of its constituent elements—Joule's law.¹ H. Kopp (1865) expressed the same idea by saying that each element has the same atomic heat in compounds as it has in the free state. This means that if 6.4 be the atomic heat of each element, and the compound contains n atoms, the molecular heat of the compound will be approximately $6.4n$. Otherwise expressed, the quotient obtained by dividing the molecular heat of a compound by the number of elementary atoms in one molecule is approximately equal to 6.4. There is obviously a difficulty with the carbonates. This is possibly connected with the difficulty previously found for carbon. Further, if the atomic heats of all but one of the elements in a compound be known, the unknown atomic heat can be computed; thus, the atomic heat of chlorine in lead chloride is $\frac{1}{2}(18.3 - 6.4) = 5.9$. A comparison of the results of experiment with calculations based upon Neumann's and Joule's laws is indicated in the following table:

TABLE XLII.—MOLECULAR HEATS OF SOLIDS.

Compound.	Formula.	Sp. heat.	Mol. weight.	Molecular heat.	
				Observed.	Calculated.
Mercuric chloride .	HgCl ₂	0.0689	270.92	18.67	19.2
Mercuric iodide .	HgI ₂	0.0420	453.84	19.06	19.2
Mercurous chloride .	HgCl	0.0520	235.46	12.25	12.8
Mercurous iodide .	HgI	0.0385	326.92	12.91	12.9

There are many discrepancies, as might be expected, and for the reasons stated in connection with the atomic heats of the elements. This is illustrated by the fact that E. Donath (1878) deduced 120 for the atomic weight of uranium from the specific heat of uranosouranic oxide, whereas C. Zimmermann's value for the specific heat gave twice Donath's value, viz. 240.

EXAMPLES.—(1) Calculate the specific heat of solid oxygen given the specific heat of potassium chlorate, KClO₃, 0.194; and that of potassium chloride 0.171. Here the molecular heat of potassium chlorate (molecular weight \times specific heat) is 25.7; and of potassium chloride, 12.8. The difference $25.7 - 12.8 = 12.9$ repre-

¹ This rule is sometimes called after A. C. Woestyn, 1848; and after H. Kopp, 1864-65.

sents the molecular heat of O_2 , hence the atomic heat of oxygen will be $\frac{1}{2}$ of $12.9 = 4.3$. By definition, atomic weight \times specific heat of solid = atomic heat of solid = 4.3. Hence, the specific heat of solid oxygen will be $4.3 \div 16 = 0.27$.

(2) The specific heat of silver chloride is 0.0911, and that of silver, 0.057; assuming the atomic weight of silver to be 107.9, what is the specific heat of solid chlorine? The molecular heat of silver chloride is 13.1, and the atomic heat of silver is 6.2. The difference $13.1 - 6.2 = 6.9$ represents the atomic heat of solid chlorine. The specific heat of solid chlorine is therefore $6.9 \div 107.9 = 0.064$.

The molecular heats can be employed to rectify the atomic weights of elements which do not form volatile compounds. Thus, the analyses of mercurous and of mercuric salts indicate that the atomic weight of mercury may be 100, 200, . . . If the atomic weight be 100, the formula of mercurous chloride will be Hg_2Cl , and of mercuric chloride, $HgCl$; while if the atomic weight be 200, the formulæ will be those indicated in the above table.

EXAMPLES.—(1) The analysis of barium chloride furnishes 35.5 parts of chlorine per 68.7 parts of barium. The specific heat of barium is 0.0465. What is the atomic weight of barium, when the atomic weight of chlorine is 35.5? The formula of barium chloride may be written Ba_xCl , where x is to be determined. The atomic weight of barium, by Dulong and Petit's rule, will be of the order $6.4 \div 0.0465 = 137$. Taking 35.5 as the atomic weight of chlorine, the fraction x must be of the order $68.7 \div 137 = \frac{1}{2}$. Hence the formula of barium chloride is Ba_2Cl , that is, $BaCl_2$, or some multiple of this. Hence the atomic weight of barium (chlorine, 35.5) must be $2 \times 68.7 = 137.4$.

(2) The percentage composition of platinum chloride is: Platinum, 57.7; chlorine, 42.3. The specific heat of platinum is 0.0324. What is the atomic weight of platinum? Hint, see (2) in the last but one set of examples. The ratio of the constituent elements is as 48.6 : 35.5; the atomic weight is of the order 197.5; the ratio x is nearly $\frac{1}{2}$; and hence the formula of the chloride is $PtCl_4$ or some multiple of this. Hence, assuming the atomic weight of chlorine is 35.5, the atomic weight of platinum will be $4 \times 48.6 = 194.4$.

Gases and liquids.—The molecular heats of gases vary according as the specific heats have been determined at constant volume or constant pressure. The ratio of the two molecular heats of gases corresponds with the values determined for the ratio of the two specific heats. The molecular heats of monatomic gases at constant volume approximate to 3; of diatomic gases, to 5; of triatomic gases, $7\frac{1}{2}$, etc. The molecular heats of liquids are generally greater than of the corresponding vapours.

The meaning of Dulong and Petit's rule.—The fact that the atomic heats of all elements are approximately the same, led Dulong and Petit to infer that "the thermal capacity of all atoms is the same." This means that every atom—no matter of what kind, requires the same amount of heat to raise its temperature 1° . Joule's rule means that each elementary atom retains the same capacity for heat when it is combined as it had when free. The number and kind of other atoms present and their mode of combination seem to have no influence on the numerical value of this property. It must be added that we can form no real conception of the "temperature of an atom" or of the "temperature of a molecule." All our conceptions are based on the properties of atoms and of molecules *en masse*. It is often stated that at the "absolute zero" of temperature, -273° , atomic motion must cease. This is a mere assertion, of no intrinsic value, and probably wrong. The statement *might* be true of the translatory motion of atom or molecule. According to the kinetic theory, temperature is proportional to the kinetic energy of the molecules;

and accordingly, as indicated previously, Dulong and Petit's rule points to a similar relation. The observations of Neumann and Joule indicate that the constituent atoms of a solid compound behave as if the solid were a mechanical mixture of its component atoms, and each atom were free to vibrate independently of the others (see p. 147).

In 1871 L. Boltzmann showed that the kinetic and potential energies of the molecules of a monatomic solid vibrating about a position of equilibrium are equal in magnitude, and that the total energy of the vibrating solid is shared equally between the average kinetic and potential energies, and is twice the value of either alone. If a monatomic gas be in contact with a solid with monatomic molecules, the bombardment of the gaseous molecules will produce a state of thermal equilibrium when the mean kinetic energy of the vibrating atoms of the solid is equal to the mean kinetic energy of the translatory motions of the molecules of the gas. It is shown on p. 692 that the kinetic energy due to the molecular motions is $\frac{3}{2}RT$, and therefore the total energy—kinetic and potential—is $2 \times \frac{3}{2}RT$, or $3RT$, where R is nearly equivalent to two calories per gram atom per degree. Accordingly, the atomic heat of the solid will be $3R = 6$ nearly. Thus Boltzmann arrived at a reasoned explanation of Dulong and Petit's law which had been derived empirically at the beginning of the century. Still later, A. Einstein, in 1907, extended to heat an earlier hypothesis of Planck (1906) on the radiation of light. The evolution and absorption of energy are not supposed to be continuous processes, but they are assumed to take place *per saltum* in quantities ϵ , 2ϵ , 3ϵ , . . . but not in intermediate amounts $\frac{1}{2}\epsilon$, $\frac{1}{3}\epsilon$, . . . The mathematical treatment of this hypothesis gives the expression :

$$\text{Atomic heat} = 3R \frac{u^2 \epsilon^2}{(\epsilon^u - 1)^2}$$

where u is written in place of $\beta\nu/T$, and β is written in place of h/k , where k is the atomic gas constant and is equivalent to the ordinary gas constant R divided by the number of atoms per gram atom of gas (approximately 6.06×10^{23}); h is a fundamental constant numerically equal to 6.62×10^{-27} , such that $\epsilon = h\nu$, where ν represents the frequency of atomic vibrations which can be determined from spectrometric observations in the infrared. The magnitude ϵ is called a quantum of energy; and the fundamental hypothesis, the quanta theory of energy. When T is large, the correction is small, and the atomic heat reduces to Boltzmann's $3R$. Einstein's formula connecting the atomic heats of solids with their vibration frequencies was very promising, since it gave results in close agreement with observations on the variations of the atomic heats of silver, diamond, etc., over a range of temperature from about -238° to about 900° . W. Nernst and F. A. Lindemann (1911), P. Debye (1913) and others have since improved the fundamental assumptions so much that J. H. Jeans (1914) could write: "The application of the quantum theory to the explanation of low temperature specific heats, from its complete naturalness, and from its agreement with experiment, seems destined to be final."

Questions.

1. How may the composition of air be determined? Would the composition of the air be represented either by the formula N_2O or $N_4 + O$? Give reasons for your answer.—*Cambridge Senior Locals.*

2. State the law of Dulong and Petit and explain its application. Discuss its limitations to the determination of atomic weights.—*Coll. of Preceptors.*

3. Classify in some logical way the various components of the air. Give proof that the air is a mechanical mixture. Explain how nitric acid is manufactured from certain components of the air, and tell why this process promises to be of very great importance.—*Princeton Univ., U.S.A.*

4. Show: (a) How a mixture of chlorine and hydrochloric acid may be freed from the latter; (b) How hydrochloric acid may be freed from moisture; (c) How ammonia may be freed from moisture; (d) How a mixture of nitrogen and oxygen may be freed from the latter; (e) How a mixture of nitrogen, hydrogen, and hydrochloric acid may be freed from the latter two.—*Amherst Coll., U.S.A.*

5. A given volume of air is left in contact with lime (calcium oxide); a second volume of air is shaken up with an acidified solution of ferrous sulphate. Carefully describe the changes which occur in the air and in the reagent in each case.—*Univ. North Wales.*

6. When the dross or skimmings from molten aluminium are allowed to cool, and then treated with water, the smell of ammonia is often in evidence. Suggest an explanation based on the known properties of aluminium.

7. The specific heat of iron at temperatures up to 100° is 0.1138 , but at 1000° it has been found to be 0.218 . If the law of Dulong and Petit holds good at high temperatures, what would the atomic weight of iron be under such altered conditions?—*Science and Art Dept.*

8. Name the gases which form the atmosphere, and describe the part which each plays in connection with animal and vegetable life.—*Tasmania Univ.*

9. Give a succinct account of the early experiments of Cavendish on the composition of air, and of the further investigations of Lord Rayleigh and Prof. Ramsay which led to the discovery of a new constituent of the atmosphere.—*Science and Art Dept.*

10. An oxide of a metal contains 48.0 per cent. of oxygen. What is the exact equivalent of the metal? If the specific heat of the metal was found to be 0.123 , what is the probable atomic weight and valency of the metal, and what would be the probable nature of the oxide?—*Inter. B.Sc., Manchester.*

11. What facts, with regard to carbon, boron, and silicon help to explain their anomalous behaviour as regards the law of Dulong and Petit? Give illustrations of the value of this law in determining atomic weights.—*London Univ.*

CHAPTER XXX

PHOSPHORUS

§ 1. Phosphorus—Occurrence and Manufacture.

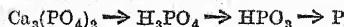
Atomic weight, $P = 31$; molecular weight, $P_4 = 124$. Ter- and quinquevalent. Melting point of yellow, 44° ; boiling point, 290° . Specific gravity at 10° , 1.83; vapour density, 127.1 ($H_2 = 2$); 4.42 (air = 1).

Bone-ash.—The animal skeleton is mainly composed of bone. Bones contain non-combustible mineral matters, and combustible organic matters. Fatty organic matter can be extracted by digesting the bones with solvents like benzene, carbon disulphide, ether, chloroform, etc., which dissolve the fat, and leave behind "degreated bones"; the gelatinous matters—glue—can be extracted by digesting the bones with water heated under pressure—"degelatinized bones" remain behind. The degelatinized bones still contain combustible carbonaceous matters, and if they be heated in iron retorts, out of contact with air, gaseous and liquid products distil over, and a residue of "animal charcoal" remains in the retort. Dippel's bone oil occurs among the liquid products of the distillation. A similar result is obtained if degreated or raw bones be heated in the retorts. If animal charcoal, or degreated, degelatinized, or raw bones be heated in air, the organic matters burn, and *bone-ash* remains. Bone-ash is used in the manufacture of manures, phosphorus, English porcelain, cupels, baking powder, etc.

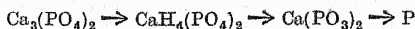
The manufacture of phosphorus—retort process.—If powdered bone-ash be intimately mixed with sulphuric acid, white insoluble calcium sulphate is formed. It is best to add just sufficient sulphuric acid to transform all the lime into calcium sulphate. The insoluble matters are separated by decantation and filtration and washed. The clear filtrate is evaporated to a syrupy liquid; mixed with about one-fourth its weight of coke or charcoal, and dried by heating in cast-iron pots. The dried mixture is then heated to redness in fireclay retorts with their necks dipping under the surface of the water. A mixture of hydrogen and carbon monoxide gases escapes, and crude phosphorus condenses as a dark brown solid. From this experiment it is inferred that bone-ash is a compound of phosphorus and lime.

The chemistry of the process is somewhat as follows: Bone-ash is a more or less impure form of calcium phosphate, $Ca_3(PO_4)_2$. When this is treated with sulphuric acid, calcium sulphate, $CaSO_4$, separates, and either normal phosphoric acid, $Ca_3(PO_4)_2 + 3H_2SO_4 = 3CaSO_4 + 2H_3PO_4$, or calcium superphosphate, $CaH_4(PO_4)_2$, remains in solution. When the evaporated solution is heated, metaphosphoric acid, HPO_3 , is formed: $H_3PO_4 = H_2O + HPO_3$, in the one case, and in the other case calcium metaphosphate, $Ca(PO_3)_2$, is produced; either of these compounds, in turn, when heated with carbon, is reduced to the element phosphorus: $2HPO_3 + 6C = 6CO + H_2 + 2P$. With calcium metaphos-

phate, silica and carbon are employed, and calcium silicate is a product of the reaction: $\text{Ca}(\text{PO}_3)_2 + 5\text{C} + \text{SiO}_2 = \text{CaSiO}_3 + 5\text{CO} + 2\text{P}$. The one scheme may be represented:



and the other scheme:



The manufacture of phosphorus—electrical process.—If bone-ash be mixed with very finely divided silica and coke, and heated in fireclay retorts to a high temperature, a similar product is obtained. The yield, however, is not so good unless the temperature is inconveniently high. If the mixture be heated in an electrical furnace fitted with carbon rods for conducting the electric current as illustrated in Fig. 234, the production of the high temperature presents no difficulty, and phosphorus distils over. Liquid slag is periodically tapped and run from the bottom of the furnace *D*, and a new charge introduced so that the process is continuous. The charge is fed into the hopper *A*, and thence passes into the chamber *B*, and to the conveyer *C*, which works something like an Archimedeian screw, and carries the charge to the furnace. At the beginning of the operation an alternating current is sent through a pair of thin carbon "electrodes," not shown in the diagram, until the furnace is hot. When heated, the resistance of the furnace is reduced, and a current is sent through the electrodes *E*, and the thin electrodes are withdrawn. The phosphorus vapours and gases escape through a tube *G*. To prevent misunderstanding, it may be necessary to point out that the electric current does its work by raising the temperature of the mass, not by electrolysis.

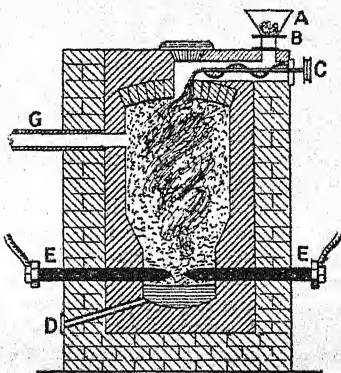


FIG. 234.—Electrical Furnace for the Manufacture of Phosphorus.

The chemistry of the process is somewhat as follows: When calcium phosphate is heated with finely divided silica, SiO_2 , calcium silicate CaSiO_3 , and phosphoric oxide are produced: $\text{Ca}_3(\text{PO}_4)_2 + 3\text{SiO}_2 = 3\text{CaSiO}_3 + \text{P}_2\text{O}_5$. The latter is reduced by the carbon: $\text{P}_2\text{O}_5 + 5\text{C} = 5\text{CO} + 2\text{P}$. The addition of a flux keeps the calcium silicate in a fluid condition and enables it to be tapped from the furnace as a molten slag. The carbon probably accelerates the rate of decomposition of the phosphate by the silica because the reaction progresses more quickly at a lower temperature in the presence of carbon than when carbon is absent. Mineral phosphates (*e.g.* sombrerite) are also used in place of bone-ash for the preparation of phosphorus.

The purification of phosphorus.—The crude phosphorus obtained by the processes of manufacture just described, contains carbon and other impurities. The methods of purification used by many manufacturers are guarded as trade secrets. According to the text-books, phosphorus is purified by redistillation from iron retorts, or by warming the phosphorus

with a mixture of sulphuric acid and potassium dichromate so as to oxidize some of the impurities; other impurities rise to the surface as a scum, and leave a layer of clear and colourless phosphorus at the bottom of the tank. The molten phosphorus is then allowed to flow into a tube of half-inch bore of such a length that the phosphorus has time to cool before it reaches the end of the tube. The phosphorus is drawn from the tube under water, and cut into sticks about $7\frac{1}{2}$ inches long. Nine sticks weigh about a pound.

The occurrence of phosphorus.—Phosphorus, of course, does not occur free in nature, because it is so very readily oxidized in contact with air. It is, however, rather widely distributed in combination with oxygen—as earthy phosphates—in such minerals as *sombrerite*, *phosphorite*, *coprolites*, and “phosphate rock” (of South Carolina, Florida, and Tennessee), all of which are more or less impure calcium phosphates $\text{Ca}_3(\text{PO}_4)_2$. The commercial value of these substances is determined by the amount of phosphorus they contain. It also occurs in *chlor-apatite*— $3\text{Ca}_3(\text{PO}_4)_2 \cdot \text{CaCl}_2$; *fluor-apatite*— $3\text{Ca}_3(\text{PO}_4)_2 \cdot \text{CaF}_2$; *vivianite*— $\text{Fe}_3(\text{PO}_4)_2 \cdot \text{H}_2\text{O}$; *wavellite*— $4\text{AlPO}_4 \cdot 2\text{Al}(\text{OH})_3 \cdot 9\text{H}_2\text{O}$. Some native phosphates are valued for the rare earths associated with the phosphoric acid—e.g. *monazite*, and some phosphates are present in certain gems—*turquoise*, *lazulite*, etc.

The phosphorus cycle in nature.—Small quantities of phosphates are found in granitic rocks. By the weathering and decay of these rocks, the combined phosphorus ultimately finds its way into the soil, spring water, and the sea. All fertile soils contain phosphorus. A ton of average fertile soil contains an average of about a pound of phosphorus—phosphorus is needed to build up certain essential parts of the tissue of growing vegetation. Plants require phosphates from their earliest life, and seeds contain a plentiful supply for the germination of the embryo. The equivalent of one pound of phosphorus is said to be present in about 100 lbs. of corn, and in about 1200 lbs. of fodder. Animals feeding upon plants or upon herbivorous animals concentrate the phosphorus in their bones and tissue. The bones of an adult man furnish about $4\frac{1}{2}$ lbs. of normal calcium phosphate. A normal adult excretes the equivalent of 3–4 grms. of phosphoric acid daily. This is derived from vegetable foods, and partly from the waste of muscular and nervous tissue which involves a decomposition of the phosphorus compounds. The products of decomposition are carried by the blood to the kidneys, and there excreted with the urine—chiefly as sodium ammonium phosphate. Physiologists have claimed that there is a relation between the amount of phosphorus compounds discharged from the system and the activity of the brain. Hence, it is inferred that phosphorus is a metabolic product of the activity of the brain, and that phosphate foods are therefore needed for the nutrition of the brain. This idea has crystallized in the phrase *Ohne phosphor, kein Gedanke* (without phosphorus no thought). The phosphorus discharged by animals finds its way back to the soil, or rather into the sewage and finally into the sea. The amount returned from the sea as edible fish is insignificant in comparison with that which was formerly drained into the sea as sewage from large towns. The phosphorus which escapes into the sea accumulates at the bottom, and there it must lie unproductive for countless ages until it is again able to take part in Nature's great cycle of changes. Processes of arresting the phosphoratic sewage have attracted

much attention. The constant growth of crops by the farmer impoverishes the soil, and phosphatic manures are needed to make good the loss. The metabolic origin of the phosphorus in the large deposits of guano—the excrement of seabirds—on the islands off the Peruvian coast, and a number of islands in the South and the Caribbean Seas, will now be obvious.

History.—It is generally supposed that phosphorus was accidentally discovered by Brand, an alchemist in Hamburg, between 1668 and 1669, while distilling a mixture of sand and concentrated urine, during his quest for a substance which would turn the base metals into gold. We are told that Brand's secret was sold to D. Krafft for about £30. It is possible that the secret was communicated to, or else independently discovered by, J. Kunckel about 1678. A couple of years later, 1680, R. Boyle, stimulated by the rumours of the process of preparation by Brand, also found, by the aid of his assistant A. G. Hanckwitz, a method of preparing phosphorus similar to that employed by Brand. In 1769, G. Gahn showed that calcium phosphate occurs in bones, and K. W. Scheele, in 1771, obtained phosphorus from bone-ash.

In the eighteenth century, the term "phosphorus" was applied to substances like commercial barium and calcium sulphides which, after exposure to a bright light, phosphoresce (appear luminous) in the dark. The word "phosphorus" is derived from the Greek— $\phi\omega\varsigma$ (phos), light; $\phi\epsilon\rho\omega$ (phero), I carry. The term "phosphorus" was naturally applied to the substance discovered by Brand, because it, too, glowed in the dark. Terms like "Brand's phosphorus," "Kunckel's phosphorus," "Boyle's phosphorus," "English phosphorus," "phosphorus mirabilis," and "phosphorus igneous," were used by the early writers on chemistry. Later the term "phosphorus" was restricted to the element phosphorus now under discussion. A. L. Lavoisier proved that phosphorus is an element, and he investigated some of the products formed when phosphorus is burned in air.

§ 2. The Allotropic Modifications of Phosphorus.

It is generally stated that there are two well-defined allotropic modifications of phosphorus—yellow and red. Many other so-called allotropic forms have been described, but there is some uncertainty whether the discoverers have not been deceived by the effect of impurities on the properties of phosphorus. Some consider the alleged "black phosphorus" to be a mixture of phosphorus with some metallic phosphide. The allotropy of phosphorus is more complex than is the case with sulphur, and it is not yet possible to give the conditions under which even the better known allotropic forms are stable.

The properties of yellow phosphorus.—This form of phosphorus has also been called "octahedral," "common," "colourless," "white," and "non-metallic" phosphorus. It is a translucent, almost colourless solid which soon becomes coated with a white opaque crust. If exposed to light, yellow phosphorus rapidly darkens in colour. At 0° it is hard and brittle with a crystalline fracture; at ordinary temperatures it is soft enough to be cut with a knife. It melts at 44° under atmospheric pressure, and at 52·8° under a pressure of 300 atmospheres. When molten phosphorus is cooled, the temperature can be reduced much below the melting

point without solidification, owing to the effects of undercooling. Phosphorus boils at 278.3° (762 mm.), and at 165° under a pressure of 120 mm., but it vaporizes at a much lower temperature. For instance, when phosphorus is heated to 40° in the lower part of a flask in an atmosphere of carbon dioxide, crystals of phosphorus sublime on to the upper part of the flask. Yellow phosphorus also sublimates at ordinary temperatures *in vacuo* when exposed to the light. Yellow phosphorus ignites at about 30° in air, saturated with moisture; the inflammation temperature is higher the dryer the air. When phosphorus burns in air, dense white clouds of phosphorus pentoxide, P_2O_5 , are formed. If, however, the phosphorus and oxygen be perfectly dried by exposure to the desiccating action of phosphorus pentoxide, phosphorus may be melted and even distilled in oxygen with very little oxidation.

The inflammation temperature is so low that the heat of the body suffices to raise the temperature of the phosphorus above its kindling temperature, and hence phosphorus should always be "handled" with the forceps, never with the bare fingers unless under water. Burns produced by phosphorus are very painful, and heal very slowly. The fumes of phosphorus are poisonous. Persons constantly exposed to their action are liable to suffer from caries (rotting) of the bones of the jaw and nose—"fossy jaw." Phosphorus itself is poisonous; it is stated that a 0.1 gram dose is fatal to man.

On account of its inflammability, phosphorus is usually preserved under water in which it is almost insoluble—one part of phosphorus dissolves in 300,000 parts of water, but it readily dissolves in carbon disulphide, sulphur chloride, phosphorus trichloride, etc.; and it is also soluble in chloroform, benzene (1.513 per cent. at 0° , and 10.03 per cent. at 81°), turpentine, alcohol, ether (0.43 per cent. at 0° , 2.00 per cent. at 35°) almond oil, (1.25 per cent.) etc. Octahedral crystals of yellow phosphorus are obtained when the solution in carbon disulphide is allowed to evaporate. If a piece of blotting paper be soaked with a solution of phosphorus in carbon disulphide, and the solvent be allowed to evaporate, the finely divided phosphorus which remains behind ignites spontaneously in air owing to its rapid oxidation. If a little phosphorus be placed in water, and the water is boiled, phosphorus volatilizes with the steam; if the vapour be passed through a condenser with the apparatus in the dark, a luminous ring appears where the steam condenses. If much phosphorus be present, globules of phosphorus will collect in the receiver. This is the principle of E. Mitscherlich's process for the detection of phosphorus.

The oxidization of phosphorus.—When phosphorus is exposed to ordinary air in the dark, it emits a pale greenish light, and gives off white fumes with an unpleasant garlic-like smell. The fumes of phosphorus in moist air are mainly P_4O_6 , and the glowing or phosphorescence of phosphorus appears to be an effect of oxidation—no oxidation, no glow. In oxygen at atmospheric pressure, phosphorus pentoxide, P_2O_5 , is the sole solid end-product of the reaction; but under reduced pressure (18–20 mm.) phosphorus trioxide, P_2O_3 , and a yellow-product (*vide infra*) are formed. When phosphorus vapour is oxidized at ordinary temperatures by the gradual admission of oxygen, a dark red-product is formed in which the ratio of oxygen to phosphorus increases as the rate of addition of oxygen increases. V. Kohlschütter and A. Frumkin consider the red product

to be a kind of intermediate "compound" of red phosphorus with a phosphorus oxide. Phosphorus does not glow at temperatures much below 0° , and not in inert gases—carbon dioxide, nitrogen, etc. Phosphorus does not glow in pure oxygen below 15° , but it does phosphoresce if the oxygen be slightly rarefied, or diluted with an inert gas; and the glow in air is stopped if the air be compressed. Thus a certain critical pressure of the oxygen is needed to start the phosphorescence. The glow in air is stopped if traces of gases which destroy ozone be present—e.g. hydrogen sulphide, ethylene, turpentine, etc. Hence ozone appears somehow to be connected with the phosphorescence. Ozone is one product of the oxidation of phosphorus in moist air.

✓ The formation of ozone during the oxidation of phosphorus takes place in such a way that the amount of ozone formed is proportional to the amount of phosphorus oxidized, such that the *atomic* ratio: Phosphorus oxidized: Ozone formed = 1:0.5. It is not clear what actually takes place. A great number of more or less plausible hypotheses have been suggested, but a crucial experiment has not been yet devised. One hypothesis assumes that atoms of oxygen alone take part in the oxidation, and that some of the molecules of oxygen are dissociated into two atoms, each atom acquires at the same time an electric charge of opposite sign: $O_2 = O^{+} \times O^{-}$. The phosphorus combines with the atoms with one kind of electric charge; and the atoms of oxygen with a charge of opposite sign form ozone. The dissociation of the oxygen molecule is not a consequence of the oxidation, but antecedent to it. Another hypothesis assumes that each atom of phosphorus is first oxidized by direct union with a whole molecule of oxygen forming, say, P_2O_3 . This then decomposes, giving atomic oxygen: $P_2O_3 = P_2O + O$. The atomic oxygen then unites with ordinary oxygen to form ozone. According to this view of the reaction, the oxygen molecule is dissociated during the oxidation of phosphorus.

The formation of red phosphorus.—When yellow phosphorus is exposed to air and light; or when phosphorus is heated to about 250° in an inert gas or *in vacuo*; or when phosphorus is exposed to an electric discharge—silent or spark—it soon passes into a chocolate-red coloured allotropic modification which sometimes has an iron-grey metallic lustre. This variety is called red phosphorus. The conversion of yellow into red phosphorus is conveniently illustrated by sealing up a small piece of yellow phosphorus in a glass tube, A, Fig. 235, and hanging the tube in the vapour of boiling diphenylamine (boiling point 310°). The vapour of the diphenylamine condenses in the long neck, and runs back into the flask. The yellow phosphorus will soon be converted into the red variety. The change proceeds with almost explosive violence when the phosphorus is heated under pressure to about 350° , that is about 60° above the boiling point of phosphorus; at 300° the change is moderately fast. The speed of the transformation can be accelerated by the addition of a trace of iodine. In the presence of this catalytic agent, the change is fairly fast at 200° .

The manufacture of red phosphorus.—Red phosphorus is made commercially by heating yellow phosphorus in a glass or porcelain vessel

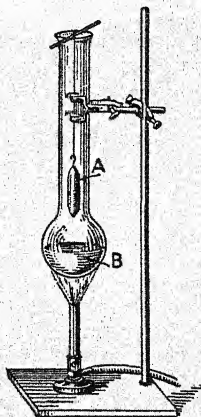


FIG. 235.—The Preparation of Red Phosphorus (L. Meyer).

embedded in sand placed in a large iron pan. The vessel containing the phosphorus is covered with an air-tight lid, and provided with a safety tube dipping in water in case the pressure inside the closed vessel becomes too great. When most of the air has been expelled, the safety tube is closed, and the phosphorus is heated until it is converted into the red variety. The product of the action is ground with water, and boiled with sodium hydroxide solution so as to remove the unaltered yellow phosphorus. The residual red phosphorus is then washed with water, and dried, or preserved wet with water.

The transition point: red to yellow phosphorus.—Red phosphorus is the stable form at ordinary temperatures, and yellow phosphorus the unstable modification. The reason the yellow phosphorus does not pass into the red form at ordinary temperatures is due to the extremely slow velocity of the change. By the distillation of red phosphorus at 290° , yellow phosphorus is obtained, and at ordinary pressures the transition point of the stable red into the unstable yellow is masked by the vaporization of the phosphorus. However, when red phosphorus is heated

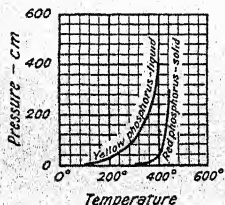


FIG. 236.—Vapour Pressure Curves of Phosphorus.

under pressure in capillary tubes so as to prevent distillation, it forms a yellow fluid at 610° , and red particles begin to separate from the cooling solution at 580° . At 570° the mass turns red. Red phosphorus melts at 589.5° under 43 atm. pressure. The vapour pressure of yellow phosphorus is greater than the red; the curves cannot be carried above 400° because the yellow phosphorus passes into the red variety so quickly at about 400° . The two vapour pressure curves are illustrated in Fig. 236. Both curves converge towards the melting point as is the case with allotropic modifications of other elements. It is inferred that both varieties would be in equilibrium with the vapour phase at the melting point—near 600° —were it not for the disturbing phenomenon just indicated.

Solubility of red phosphorus.—Red phosphorus is practically insoluble in the ordinary solvents—*e.g.* turpentine—which dissolve yellow phosphorus. Red phosphorus is soluble in phosphorus tribromide, and in alcoholic potash. If the solution in phosphorus tribromide be exposed to light, or boiled for some time, what is supposed to be a mixture of red phosphorus and phosphorus tribromide separates as a scarlet red powder—Schenk's scarlet phosphorus. This resembles red phosphorus in many of its properties, but is rather more chemically active, as might be expected from its extremely fine state of subdivision. This variety of red phosphorus has not been prepared free from the solvent in which it is formed; consequently, the so-called scarlet phosphorus appears to be a mixture of red phosphorus with phosphorus tribromide; that deposited at 30° by exposure to light retains about 47 per cent. of the solvent; at 140° , about 36 per cent.; at 185° , about 27 per cent.; and at 218° , about 24 per cent. When the red-coloured solution of red phosphorus in alcoholic potash is acidified, a red powder is precipitated. Some believe that this is nothing but impure red phosphorus; others believe that it is an impure phosphorus suboxide— P_4O . E. Jungfleisch (1907) said that if the yellow

substance formed when phosphorus is oxidized in oxygen under reduced pressure is dissolved in alcoholic potash, and precipitated by hydrochloric acid, the same suboxide is formed. The analysis of the alleged suboxide is difficult because of the impurities associated with the precipitate, and hence the question has not been definitely answered.

The crystallization of red and yellow phosphorus.—Crystals of yellow phosphorus are obtained by the slow evaporation of a solution of phosphorus in carbon disulphide; or by sublimation *in vacuo* or in an atmosphere of an inert gas; they belong to the cubic system. Small rhombohedral crystals of red phosphorus are formed by heating ordinary red phosphorus under pressure at 580°, or in a sealed tube with metallic lead at about 500° for about eighteen hours. The lead dissolves the phosphorus at the high temperature, and rejects the dissolved phosphorus on cooling in the form of rhombohedral crystals. The lead can be dissolved away by means of dilute nitric acid followed by boiling with concentrated hydrochloric acid; the residue has a specific gravity varying from 2.34 to 2.39. This dense form of red phosphorus is sometimes called Hittorf's phosphorus, or metallic phosphorus, or "rhombohedral" phosphorus. The term "metallic" appears to be a misnomer, for it does not conduct electricity. The specific gravity of red phosphorus varies with its method of preparation from 2.05 to the maximum 2.34 for Hittorf's phosphorus.

Since the amount of heat developed by the combustion of red phosphorus is not constant, it is possible that ordinary red phosphorus is a solution of colourless phosphorus in the rhombohedral variety; or else a mixture of colloidal and "metallic" phosphorus. E. Cohen and J. Olie (1909) consider red phosphorus to be an isomorphous mixture or a solid solution of yellow and metallic phosphorus in dynamic equilibrium: $P_{\text{red}} \rightleftharpoons P_{\text{yellow}}$. A. Smits and H. L. de Leeuw (1911) consider yellow, red, and metallic phosphorus to be three solids which can form solid solutions with one another.

Properties of red and yellow phosphorus.—The following table gives a summary of the leading differences in the two varieties of phosphorus:

TABLE XLIII.—COMPARISON OF THE PROPERTIES OF RED AND YELLOW PHOSPHORUS.

Property.	Red phosphorus.	Yellow phosphorus.
Colour	Reddish-violet	Almost colourless
Crystalline form	Rhombohedral system	Cubic system
Smell, etc.	Tasteless, odourless	Garlic-like smell ¹
Exposed to air	No phosphorescence no oxidation	Phosphorescence and oxidation
Melting point	500° to 600°	44°
Physiological action . .	Non-poisonous	Poisonous
Specific gravity	2.05 to 2.39	1.83 to 1.85
Specific heat	0.170	0.189
Action carbon disulphide	Insoluble	Soluble
Ignition temperature	280°	30°
Electric current	Feeble conductor	Very feeble conductor
Hot sodium hydroxide . .	Nil	Action
Chlorine gas	Fires if heated	Fires spontaneously
Heat of combustion (Cals.)	363.4	370.8

¹ When oxidizing.

P. W. Bridgman's study of the effect of high pressures on phosphorus furnished him with an hexagonal variety of yellow phosphorus which passes into the ordinary cubic form at the transition point -76.9° ; and a variety of black phosphorus formed at 200° under pressures from 12,000 to 13,000 kilograms per sq. cm.

Chemical properties.—Red phosphorus is less chemically active than ordinary yellow phosphorus. This agrees with the fact that the passage of the yellow to the red variety is attended by an evolution of heat: $P_{\text{yellow}} = P_{\text{red}} + 4 \text{ Cals.}$; and hence it is generally stated that red phosphorus contains less available energy than yellow phosphorus. Red phosphorus does not appear to volatilize appreciably below 200° , and only very slightly at 280° , whereas Hittorf's phosphorus volatilizes at that temperature. Red phosphorus takes fire when heated in air to about 260° ; and like yellow phosphorus, previously discussed, it burns, forming phosphorus pentoxide. Red phosphorus burns if heated in an atmosphere of chlorine, whereas ordinary phosphorus fires spontaneously in the gas. At ordinary temperatures, phosphorus reacts with sulphur, forming phosphorus sulphides; with the halogens, forming tri- or penta-halides; with hot potassium or sodium hydroxide, forming gaseous phosphorus hydrides. Concentrated nitric acid is reduced with almost explosive violence, while dilute nitric acid evolves nitrous fumes; boiling sulphuric acid is reduced to sulphur dioxide, etc.

The difference in the molecules of red and yellow phosphorus is not known. As is probably the case with monoclinic and rhombic sulphur, it may be due to a difference in the arrangement of identical molecules, or to a difference in the molecular weight, that is, to the number of atoms per molecule. Equal weights of red and yellow phosphorus when burnt separately in an excess of oxygen, give equal weights of phosphorus pentoxide, P_2O_5 .

Phosphides.—Phosphorus unites directly with many metals, forming phosphides. Many phosphides can be made by heating the finely powdered metal with phosphorus in a sealed tube. Metallic phosphides are precipitated by the action of phosphorus on solutions of some metallic salts. Thus silver phosphide, Ag_3P , is precipitated from solutions of silver nitrate; and copper phosphide, Cu_3P_2 , from solutions of copper sulphate. Copper phosphide is also formed in hexagonal steel-grey prisms by the action of phosphorus vapour carried in a stream of an indifferent gas, over hot metallic copper. Mercury phosphide, P_4Hg_3 , is formed in glistening rhombohedral prisms by heating phosphorus with mercury iodide in a sealed tube. Manganese phosphide, MnP_2 , has been made by heating mixtures of the two elements in a sealed tube; when heated to 400° in a stream of hydrogen it passes into MnP . Tin forms several phosphides when heated with phosphorus. For instance, tin phosphide, Sn_3P , is a coarsely crystalline mass melting at about 170° . Tin phosphide is also called "phosphor-tin," and it is used in the preparation of manganese bronzes—"phosphor-bronze," for example, is made by fusing together copper and phosphor-tin. The resulting alloy contains from 5 to 15 per cent. of tin, and 0.25 to 1.5 per cent. of phosphorus. This alloy is hard and firm. It is remarkably homogeneous, and is specially valuable for the manufacture of axle bearings, etc. Lead phosphide, PbP_3 , is formed as a black unstable solid which spontaneously inflames in air. It decomposes into its elements when

heated in vacuo at 400° ; it is slowly attacked by water; and is decomposed by dilute acids. The phosphides of the bivalent metals are formed when a solution of an alkali phosphide in liquid ammonia is treated with a salt of the metal. Calcium phosphide is formed when metallic calcium is heated under naphtha with phosphorus; or by passing the vapours of phosphorus over red hot lime heated in an iron tube. The red, brown, or black mass obtained in the last-named reaction is mixed with some calcium pyrophosphate, $\text{Ca}_2\text{P}_2\text{O}_7$. A series of alkali phosphides have been made by fractionally distilling the alkali metal with an excess of phosphorus in vacuo. These phosphides have the general formula M_2P_3 , and they react with water giving a solid hydride, and a little phosphine.

Atomic and molecular weight.—The atomic weight of phosphorus with reference to oxygen 16, lies somewhere between 30.91 and 31.05; the best representative value is taken to be 31. This has been determined by the analysis of silver phosphate; by the action of phosphorus on silver nitrate; and by the titration of phosphorus trichloride on silver nitrate. Vapour density determinations of volatile phosphorus compounds— PH_3 , PCl_3 , PF_5 , etc.—show that thirty-one is the smallest amount of phosphorus which enters into the composition of any one of these molecules, and hence the atomic weight is 31.

The vapour of phosphorus, according to E. Mitscherlich, has a specific gravity of 4.58 at 515° , and 4.5 at 1040° . Hence the molecular weight of phosphorus is $4.5 \times 28.75 = 129.4$. This corresponds with a molecule P_4 —graphically:—



Above 700° , the vapour density falls to a number corresponding with a dissociation of the complex molecule $\text{P}_4 \rightleftharpoons 2\text{P}_2$; at about 800° , one-hundredth of the phosphorus vapour exists as P_2 molecules; at 1000° , one-tenth; and at 1200° , one-third—all at atmospheric pressure. The depression of the freezing point of solutions of phosphorus in carbon disulphide also corresponds with the four-atom molecule. In some other solvents numbers corresponding with P_2 , P_4 , or with mixtures of P_4 and P_2 are obtained.

§ 3. Phosphorus Trihalides.

Phosphorus trichloride, PCl_3 .—This compound is made by passing chlorine, dried by sulphuric acid, over molten white phosphorus in the following manner:

A layer of sand is placed at the bottom of a retort, fitted up as illustrated in Fig. 171, without the Bunsen's burner, and a current of dry carbon dioxide, or another inert gas, is introduced. Add, say, 100 grams of yellow phosphorus—dried between filter paper, and dipped successively in alcohol and in ether—then pass a current of chlorine through the apparatus while the retort is heated with warm water. The tube delivering the chlorine should be movable, for if it be too near the phosphorus the latter gets hot and distils, forming a red crust in the upper part of the retort; while if it be too far away, the action is slow, and the excess of phosphorus forms phosphorus pentachloride by a side reaction. When the action has begun, a tongue of flame projects from the tube delivering the chlorine. The retort does not then need heating. Towards the end, when the phosphorus has all disappeared, heat the retort very gently so as to drive the trichloride into the receiver. The fumes from the exit tube must be led into a stink closet

or into a vessel containing sodium hydroxide. The product can be purified by adding, say, 2 grams of yellow phosphorus and redistilling. The object of the phosphorus is to convert any pentachloride into the trichloride. Moisture must be carefully excluded.

Phosphorus trichloride is a mobile liquid with an unpleasant smell. It boils at 74° ; fumes in air, and is hydrolyzed by water forming phosphorus and hydrochloric acids: $\text{PCl}_3 + 3\text{H}_2\text{O} = 3\text{HCl} + \text{P}(\text{OH})_3$. It can be frozen to a solid, melting at about -115° . The composition of phosphorus trichloride, and also of the other halides of phosphorus, can be determined by treating the compound with water, and determining the amounts of the different acids formed by the regular methods of analysis. The lowering of the freezing point of solution of phosphorus trichloride in phosphorus oxychloride, or in benzene, corresponds with the molecular weight PCl_3 . Its vapour density, according to J. B. Dumas, is 140 ($\text{O}_2 = 32$), thus corresponding with the molecule PCl_3 (theoretical vapour density 137.5). Phosphorus trichloride vapour mixed with hydrogen is converted into phosphorus dichloride, PCl_2 or P_2Cl_4 , by the electric discharge. The dischloride is a fuming colourless oily liquid which freezes to a white solid at -28° . It is unstable and decomposes slowly (sometimes explosively) at ordinary temperatures, furnishing a yellow solid with an analysis approximating P_2Cl_2 . An analogous solid was obtained in the unsuccessful attempt to make *phosphorus dibromide*.

Phosphorus tribromide, PBr_3 .—Phosphorus tribromide is made by gradually adding a solution of bromine in carbon disulphide to dry red phosphorus, or to a solution of phosphorus in carbon disulphide. The object of the carbon disulphide is to moderate the violence of the reaction. The solvent is afterwards distilled off, at about 46° ; and the tribromide is distilled, at about 174° —the boiling point of the tribromide. The general properties of phosphorus tribromide resemble those of the trichloride.

Phosphorus triiodide, PI_3 .—This compound is made by dissolving, say, 10 grams of phosphorus in carbon disulphide, and gradually adding the solution to a solution of 123 grams of iodine in carbon disulphide. The solvent is then distilled off. The triiodide remains behind as a reddish crystalline solid which melts at 41° , and dissociates when heated. A compound with the empirical formula: P_2I_4 —called phosphorus di-iodide—remains in the form of orange-red crystals, which melt at 110° , when a mixture of 50 grams of iodine and 4 grams of phosphorus, is melted.

Phosphorus trifluoride, PF_3 .—Phosphorus trifluoride is made by the action of copper phosphide on lead fluoride; or by allowing arsenic fluoride, AsF_3 , to drop slowly into PCl_3 with the exclusion of moisture. It is a colourless gas which can be condensed to a liquid, boiling at -95° , and frozen to a solid, melting at -160° . If a stream of electric sparks be passed through the gas, phosphorus and phosphorus pentafluoride are formed: $5\text{PF}_3 = 3\text{PF}_5 + 2\text{P}$; and the pentafluoride, in turn, is said to suffer slight decomposition: $\text{PF}_5 = \text{PF}_3 + \text{F}_2$.

§ 4. Phosphorus Pentahalides.

Phosphorus pentachloride, PCl_5 .—Phosphorus pentachloride is made by the action of an excess of chlorine upon phosphorus, or by the action of dry chlorine upon the trichloride; or by the action of sulphur chloride

on phosphorus trichloride with a trace of iodine as stimulant: $3\text{PCl}_3 + \text{S}_2\text{Cl}_2 = \text{PCl}_5 + 2\text{PSCl}_3$ —a yield of 88 per cent. of the theoretical is obtained. Since phosphorus pentachloride is a very unpleasant substance to manipulate in air, owing to the fact that it rapidly absorbs moisture, forming hydrochloric and phosphoric acids: $\text{PCl}_5 + 4\text{H}_2\text{O} = 5\text{HCl} + \text{H}_3\text{PO}_4$, it is best to make the compound in the bottle in which it is to be preserved. Fit the bottle with a three-hole stopper, Fig. 237—one hole is for the tube bringing in dry chlorine, one for the exit of the chlorine, and the third for a tap funnel by means of which phosphorus trichloride can be run into the chlorine drop by drop.

The vapour pressure of solid phosphorus pentachloride at 140° is 760 mm., and it therefore vaporizes without fusion. Phosphorus pentachloride also sublimes below 100° without fusion, but when heated under pressure it melts at about 148° . Both phosphorus trichloride and phosphorus pentachloride are valuable reagents for transforming hydroxyl compounds into the corresponding chlorides. Each OH group is displaced by an atom of the halogen. For example, water, $\text{H}-\text{OH}$, gives hydrogen chloride, $\text{H}-\text{Cl}$; nitric acid, NO_2-OH , gives nitroxyl chloride, NO_2-Cl ; sulphuric acid, $\text{SO}_2(\text{OH})_2$, gives sulphuryl chloride, SO_2Cl_2 ; alcohol, $\text{C}_2\text{H}_5\text{OH}$, gives ethyl chloride, $\text{C}_2\text{H}_5\text{Cl}$, etc.

Phosphorus pentabromide, PBr_5 .—This compound is prepared by a similar process to that used for the pentachloride—the addition of bromine to phosphorus tribromide—and its properties are similar. Phosphorus pentiodide— PI_5 —has been reported, but there is some doubt as to its real existence.

The vapour density of phosphorus pentachloride and pentabromide.—The vapour densities of these two compounds diminish with rise of temperature. Thus J. B. Dumas found for the pentachloride:

Temperature . . .	182°	200°	250°	300°
Vapour density . . .	146.6	140.0	115.2	104.8
Amount dissociated. .	42.2	48.9	83.6	98.9 per cent.

Theory for $\text{PCl}_5 = 208.5$

The vapour density remains practically constant at temperatures exceeding 300° . The colour of the vapour becomes yellowish-green at the higher temperatures, showing that free chlorine is probably present; and in the case of the pentabromide, the characteristic colour of free bromine appears. There is little room for doubt that phosphorus pentachloride dissociates into free chlorine (Cl_2 , vapour density 71) and phosphorus trichloride (PCl_3 , vapour density 137.5). Paper moistened with potassium iodide and starch shows the blue coloration characteristic of free chlorine when the pentachloride is heated to 157° or 158° . Assuming that dissociation is complete at 300° , the vapour density should be $\frac{1}{2}(137.5 + 71) = 104.3$ ($\text{H}_2 = 2$), a number very close to the observed value. The discussion in connection with the dissociation of nitrogen peroxide shows how the amount of dissociation can be determined from the vapour density at the different temperatures.

According to the principles developed in connection with the dissocia-

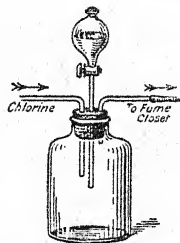


FIG. 237.
The Preparation of
Phosphorus Pen-
tachloride.

tion of iodine, the dissociation of the pentachloride will be represented by the equation: $\text{PCl}_5 \rightleftharpoons \text{PCl}_3 + \text{Cl}_2$. If C_{PCl_5} , C_{PCl_3} , and C_{Cl_2} respectively denote the concentrations of phosphorus pentachloride, phosphorus trichloride, and of chlorine, we have, for equilibrium: $kC_{\text{PCl}_5} = k'C_{\text{PCl}_3}C_{\text{Cl}_2}$. Suppose that one gram of phosphorus pentachloride be heated in a closed vessel of volume v . Let x represent the fraction dissociated at any given temperature, there will obviously be C_{PCl_5} , or $(1-x)/v$ gram-molecules of the pentachloride per unit volume; C_{PCl_3} , or x/v gram-molecules of the trichloride; and C_{Cl_2} , or x/v gram-molecules of chlorine. Hence for equilibrium, the preceding equation becomes

$$k \frac{(1-x)}{v} = k' \left(\frac{x}{v} \right)^2; \text{ or, } K = \frac{k}{k'} = \frac{x^2}{(1-x)v}$$

Suppose that at 250° 0.80 gram-molecules of the substance are dissociated, then $x = 0.8$; and $1-x = 0.2$. Hence $K = 0.64 \div 0.2 = 3.2 \div v$. One gram-molecule of a gas at 760 mm. and 0° occupies 22.3 litres, and at 250° it will occupy 42.7 litres. But on dissociation, one gram-molecule of phosphorus pentachloride becomes two gram-molecules of mixed chlorine and phosphorus trichloride. Hence if 0.8 gram-molecule of the pentachloride is dissociated, the mixed gas contains $1 + 0.8 = 1.8$ gram-molecules, so that if one gram-molecule occupies 42.7 litres, 1.8 gram-molecules will occupy $42.7 \times 1.8 = 76.9$ litres. Hence $K = 3.2 \div 76.9 = \frac{1}{24}$; or $k:k' = 1:24$. This means that the phosphorus trichloride and chlorine will unite twenty-four times as fast as the pentachloride will dissociate, supposing that each substance has unit concentration, and each action proceeds without reversion, p. 315.

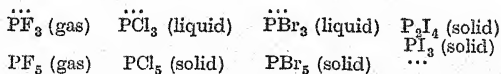
If the concentration of the phosphorus trichloride or of chlorine be augmented, it is obvious that in order to restore equilibrium, the speed of combination of the chlorine and trichloride must be augmented. This is easy to demonstrate: (1) from the theory of opposing reactions; (2) from the kinetic theory of dissociation; and (3) experimentally by sealing equal amounts of phosphorus pentabromide in two stout test-tubes, but in one of the tubes also place some phosphorus tribromide. When the two tubes are heated under the same conditions, the contents of the tube without the tribromide will have a deeper tint than the tube with the tribromide. H. Wurtz, indeed, in 1873, found that if the pentachloride be volatilized in an atmosphere of the trichloride, the vapour density is nearly normal, 206.6, between 160° and 175° . The raising of the boiling points of solutions of phosphorus pentachloride in carbon tetrachloride correspond with the formula, PCl_5 . The lowering of the freezing points of solutions of the pentachloride in benzene correspond with the molecular weight PCl_5 .

It has been argued that because phosphorus pentachloride splits up so readily into the trichloride and chlorine, two of the chlorine atoms in the pentachloride must be held less firmly by the phosphorus than are the remaining three. The inference is not valid if it be taken to mean that three of the five valences of quinquivalent phosphorus are stronger than the other two—it may simply mean that a four-atom molecule is more stable than a six-atom molecule.

Phosphorus pentafluoride, PF_5 .—This compound is made by the action of arsenic trifluoride on phosphorus pentachloride: $5\text{AsF}_3 + 3\text{PCl}_5$

$= 5\text{AsCl}_3 + 3\text{PF}_5$. It is a colourless gas which is hydrolyzed by water, forming phosphoric and hydrofluoric acids. It can be condensed to a liquid, boiling at -75° , and frozen to a solid, melting at -83° . Its vapour density is 126 ($\text{H}_2 = 2$), and its formula is therefore PF_5 (vapour density, 126), where phosphorus is undoubtedly quinquevalent.

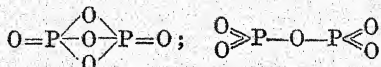
To summarize, the following well-defined halides of phosphorus have been prepared:



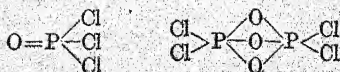
Some mixed halide salts— PF_3Cl_2 ; PF_3Br_2 ; PCl_3Br_2 ; etc.—have also been isolated.

Phosphoryl chloride, phosphorus oxychloride, POCl_3 .—This compound can be made by very carefully adding water to phosphorus pentachloride until the solid disappears: $\text{PCl}_5 + \text{H}_2\text{O} = \text{POCl}_3 + 2\text{HCl}$. It is also made by gradually adding, say, 32 grams of powdered potassium chlorate to 100 grams of phosphorus trichloride at ordinary temperatures, and then distilling the mixture. The oxychloride boils at 107.2° . It can be solidified to a colourless crystalline mass melting at -1.25° . It fumes in air, smells like the trichloride, and its aqueous solution is slowly hydrolyzed by water, forming phosphoric and hydrochloric acids. A. Besson claimed to have isolated pyrophosphoryl chloride, $\text{P}_2\text{O}_3\text{Cl}_4$, and phosphorus dioxymonochloride, PO_2Cl , from the products of the action of water on phosphoryl chloride. Pyrophosphoryl chloride has also been made by the action of water on the pentachloride, and by the action of the pentoxide on the pentachloride. The corresponding phosphoryl bromide, POBr_3 , is made in a similar manner, and boils at 190° . Phosphoryl fluoride, POF_3 , as well as the other phosphoryl compounds can be made by the action of phosphorus pentoxide, P_2O_5 , on the halogen acid: $4\text{P}_2\text{O}_5 + 6\text{HF} = 2\text{POF}_3 + 6\text{HPO}_3$. Phosphorus trichloride reacts with sulphur chloride, S_2Cl_2 , particularly if a trace of iodine be used as stimulant forming thiophosphoryl chloride, PSCl_3 , in accord with the equation $3\text{PCl}_3 + \text{S}_2\text{Cl}_2 = \text{PCl}_5 + 2\text{PSCl}_3$.

If phosphorus trichloride, PCl_3 , be oxidized at a low temperature by treatment with nitrogen peroxide, N_2O_4 , among other products, pyrophosphoryl chloride, $\text{P}_2\text{O}_3\text{Cl}_4$, is obtained as a colourless fuming liquid which boils at 210° – 215° with partial decomposition into P_2O_5 and POCl_3 . This mode of preparation is interesting because it shows that two of the oxygen atoms in phosphorus pentoxide probably behave differently from the other three. Some take this to mean that of the two formulæ for phosphorus pentoxide:



the former is to be preferred, and that the graphicæ formulæ for phosphorus oxychloride, and pyrophosphoryl chloride are respectively



Two isomeric oxychlorides of phosphorus can be predicted; in one, $\text{Cl}_2=\text{P}-\text{OCl}$, the phosphorus is tervalent, and in the other, $\text{O}=\text{P}\equiv\text{Cl}_3$, quinquivalent; but only one is known. On the other hand, two different substances, $\text{PO}(\text{C}_6\text{H}_5)_3$, are known. One (phenoxy-diphenyl phosphine) is a thick oily liquid, boiling at 265° (62 mm.); in this the phosphorus is tervalent and the compound reacts with oxygen, bromine, selenium, and sulphur, and forms additive compounds with alkyl halides. The other (triphenyl phosphine oxide) is solid, melting at 153.5° ; in this the phosphorus is quinquivalent, and it does not react with the substances just named. In both the vapour density corresponds with the empirical formula $\text{PO}(\text{C}_6\text{H}_5)_3$.

§ 5. Phosphorus Hydrides, or Hydrogen Phosphides.

Phosphine, PH_3 .—When phosphorus is heated with caustic alkaline solutions—for instance, milk of lime or a solution of potassium hydroxide

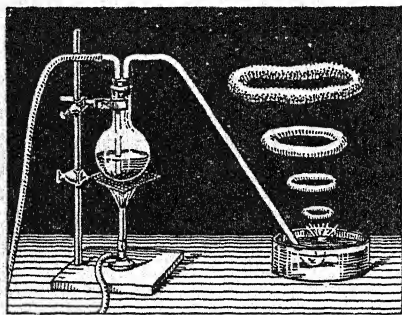


FIG. 238.—The Preparation of Phosphine.

—a gas or mixture of gases called *phosphuretted hydrogen*, is evolved which ignites spontaneously in air. This property appears to be due to the presence of a little of the so-called liquid hydrogen phosphide, P_2H_4 , as an impurity; when this impurity is removed by condensation in the cold tube, the gas does not ignite when brought in contact with air. The experiment is made by means of the apparatus illustrated in Fig. 238. A mixture of potassium hydroxide solution and yellow phosphorus is placed in the flask fitted with delivery tube, etc., as illustrated in the diagram. A current of an inert gas—coal gas, hydrogen or carbon dioxide—is first led through the apparatus to drive out the air. The mixture in the flask is then heated. The phosphorus reacts with the alkali, forming potassium hypophosphite, and gaseous phosphine, PH_3 , associated with some impurities. The mixture of gases so prepared is sometimes called “phosphuretted hydrogen.” The main reaction is represented: $3\text{KOH} + 4\text{P} + 3\text{H}_2\text{O} = 3\text{KH}_2\text{PO}_2 + \text{PH}_3$. Each bubble of gas rises to the surface of the water, and, when it comes in contact with the air, ignites with a slight explosion and burns with a brilliant flash of light, forming a vortex ring of phosphorus pentoxide, P_2O_5 . If the gas be passed through a tube immersed in a freezing mixture, it no longer inflames on contact with the air. Indeed, pure phosphine does not inflame spontaneously in air—the freezing mixture presumably removes the spontaneously inflammable liquid hydride.

The properties of phosphine.—Phosphine is a colourless gas, smells like decaying fish, liquefies about -85° , solidifies about -133° , and ignites when heated to about 100° . It burns in air, forming phosphorus pent-

oxide: $2\text{PH}_3 + 4\text{O}_2 = \text{P}_2\text{O}_5 + 3\text{H}_2\text{O} + 311.2 \text{ Cals. at constant pressure.}$
 If a mixture of oxygen and phosphine be suddenly rarefied, an explosion occurs. The phenomenon recalls the effect of rarefaction on the luminosity of phosphorus in oxygen gas. Nitric acid or chlorine when brought in contact with the gas cause inflammation. A jet of phosphine inflames and burns, forming phosphorus pentachloride when placed in chlorine gas: $\text{PH}_3 + 4\text{Cl}_2 = 3\text{HCl} + \text{PCl}_5$. Phosphine is slightly soluble in water: 100 volumes of water dissolve about 11 volumes of the gas. The aqueous solution is not alkaline like aqueous ammonia, it decomposes on exposure to light and deposits red phosphorus. The gas possesses reducing properties. When passed into solutions of copper sulphate, CuSO_4 , mercuric chloride, HgCl_2 , phosphine precipitates phosphides of the metals. It combines with ammonia, and some of the chlorides, for instance, aluminium chloride, stannic chloride, etc.

The composition of phosphine.—J. B. Dumas determined the composition of phosphine by passing a known volume of the gas over heated

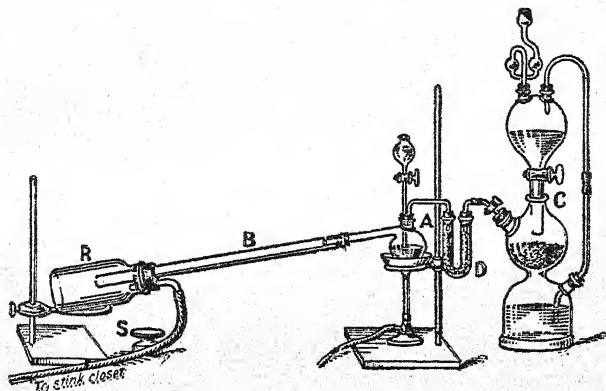
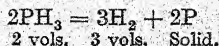


FIG. 239.—The Preparation of Phosphonium Iodide.

copper turnings—zinc, antimony, iron, and potassium have been used in place of copper. The copper forms copper phosphide. The increase in weight of the copper shows the amount of phosphorus in a given volume of the gas. The escaping hydrogen is collected and measured. Experiment shows that 34 parts of phosphine by weight give 31 parts of phosphorus and 3 parts of hydrogen. The empirical formula is therefore PH_3 . Similarly, when phosphine is decomposed by the passage of electric sparks through the gas (Fig. 185), one volume of phosphine gives $1\frac{1}{2}$ volumes of hydrogen, and red phosphorus, of negligibly small volume in comparison with the volume of the gas, is deposited on the walls of the tube. Hence two volumes of phosphine furnish three volumes of hydrogen; otherwise expressed, by Avogadro's hypothesis, two molecules of phosphine give three molecules of hydrogen:



One litre of phosphine weight 1.52 grams under normal conditions. Hence if one litre of oxygen weighs 1.429 grams the vapour density with reference to oxygen 32 is 33.9, corresponding with the formula PH_3 .

Phosphonium compounds.—When dry phosphine is brought in contact with dry hydrogen chloride, hydrogen bromide, or hydrogen iodide at high pressures, the gases unite, forming a series of so-called phosphonium compounds: $\text{PH}_3 + \text{HCl} = \text{PH}_4\text{Cl}$. The chief interest of the phosphonium compounds lies in their relationship to the ammonium compounds which are formed in a similar manner: $\text{NH}_3 + \text{HCl} = \text{NH}_4\text{Cl}$. The monad radicle, PH_4 , phosphonium, bears the same relation to phosphorus that NH_4 , ammonium, bears to nitrogen. The basic properties of phosphine are very much feebler than ammonia. Phosphonium iodide is one of the best known phosphonium compounds. It is prepared in the following manner:

Place 100 grams of phosphorus in a retort, *A*, and add an equal weight of carbon disulphide, and then 170 grams of pure iodine. Keep the retort well cooled during the mixing. Distil off the carbon disulphide by attaching a Liebig's condenser to the retort, and placing the retort in a dish of warm water. The retort is then connected with a wide tube, *B*, which may be fitted with a glass bottle, *R*, as shown in Fig. 239, and then, in the fume closet, with a wash-bottle containing water to absorb any hydrogen iodide formed by the decomposition of the phosphonium iodide. Connect the retort with a Kipp's apparatus, *C*, for generating carbon dioxide, dried by passage through a tube, *D*, charged with pumice stone soaked with sulphuric acid, and keep a slow stream of carbon dioxide passing through the apparatus all the time an experiment is in progress. Let 85 grams of water fall slowly, drop by drop, on to the residue in the retort. The heat of the reaction suffices to sublime the phosphonium iodide into the wide tube, but towards the end of the operation, the retort may be warmed. When the operation is completed, loosen the phosphonium iodide which has collected in the wide tube by means of a piece of wire, and transfer the salt to the bottle used as a receiver. The bottle is then closed with its stopper, *S*. The reaction is represented: $5\text{I} + 9\text{P} + 16\text{H}_2\text{O} = 5\text{PH}_4\text{I} + 4\text{H}_3\text{PO}_4$.

Phosphonium iodide crystallizes in large quadratic prisms with a brilliant lustre. It is an unstable salt readily dissociating into hydrogen iodide and phosphine, even at as low a temperature as 30° . The crystals can be sublimed without melting. They fume in air, and in contact with water form hydrogen phosphide and hydrogen iodide. With potassium hydroxide, fairly pure phosphine is obtained: $\text{PH}_4\text{I} + \text{KOH} = \text{KI} + \text{H}_2\text{O} + \text{PH}_3$; and with alcohol $\text{C}_2\text{H}_5\text{OH}$, a similar reaction takes place: $\text{PH}_4\text{I} + \text{C}_2\text{H}_5\text{OH} = \text{C}_2\text{H}_5\text{I} + \text{H}_2\text{O} + \text{PH}_3$. Phosphonium iodide is used as a reducing agent, and in the preparation of organic phosphines. Phosphonium chloride is dissociated at ordinary temperatures, but it can exist at 14° or under if kept under a pressure of 20 atmospheres. It must therefore be preserved in sealed tubes, and prepared by the combination of phosphine with hydrogen chloride under pressure. Similar remarks apply to phosphonium bromide, PH_4Br .

Liquid hydrogen phosphide, P_2H_4 .—When calcium phosphide, Ca_3P_2 , is treated with water, and the gas evolved passed through a spiral tube in order to condense the water, and then through a U-tube immersed in a freezing mixture (pounded ice and salt), a colourless liquid is obtained which is spontaneously inflammable when exposed to the air. The uncondensed gases which pass on are led into a trough of water in the fume closet, each bubble of gas as it comes in contact with the air may burn with a bright flash of light characteristic of crude phosphine. This is due

to the escape of some of the "liquid" phosphide. The apparatus for the experiment is shown in Fig. 240. The calcium phosphide is dropped through a wide tube into the Woulff's bottle containing water. The liquid boils at 57° to 58° (735 mm.). The empirical formula is PH_3 . The vapour density is 75.5 (theory for P_2H_4 is 66). The molecular formula is therefore P_2H_4 , and if phosphorus be trivalent, the graphic formula is probably $\text{H}_2=\text{P}-\text{P}=\text{H}_2$, analogous with hydrazine N_2H_4 .

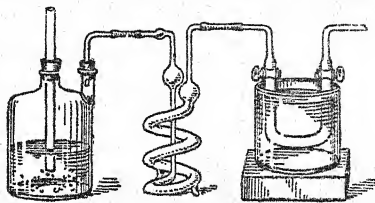


FIG. 240.—The Preparation of Liquid Hydrogen Phosphide.

Solid hydrogen phosphide P_{12}H_6 .—When liquid hydrogen phosphide is exposed to light, or heated above its boiling point, it decomposes into gaseous phosphine, PH_3 , and a solid canary-yellow phosphorus hydride with the empirical formula P_2H , thus: $5\text{P}_2\text{H}_4 = 2\text{P}_2\text{H} + 6\text{PH}_3$. This solid is insoluble in water. The same reaction takes place in the presence of hydrochloric or hydriodic acid which act as catalytic agents. The depression of the freezing point of a solution of the solid hydride in molten phosphorus corresponds with a molecular weight P_{12}H_6 . When heated to 70° in a stream of carbon dioxide, it decomposes into phosphorus and hydrogen. It takes fire when heated in air to 160° , and when heated *in vacuo* it evolves fairly pure phosphine, forming a red substance also said to be a solid orange-red phosphorus hydride with the empirical formula P_9H_2 , thus, $5\text{P}_{12}\text{H}_6 = 6\text{P}_9\text{H}_2 + 6\text{PH}_3$. Prolonged heating is said to convert this hydride into red phosphorus. If the P_9H_2 hydride be heated *in vacuo* for some hours at 80° , a further loss of phosphine occurs, and a yellow solid hydride, H_2P_5 is said to be formed. A solid hydride is also obtained among the products of the action of water on the alkali phosphides. There is some doubt about the real existence of some of these hydrides. They may be solid solutions of red phosphorus with one of the other hydrides, or with hydrogen.

§ 6. Phosphorus Sulphides.

Several compounds of sulphur and phosphorus have been reported. Most of them are made by the direct combination of heated mixtures of the two elements. The reaction between yellow phosphorus and sulphur is very violent, and red phosphorus is therefore used. The sulphur also should be coarsely granulated in order to keep down the velocity of the reaction. The mixture of sulphur and phosphorus is placed in a flask with a cork placed loosely in the neck. The mixture is heated on a sand bath until the reaction starts; the flame is then removed. This method of preparation is called Kekulé's process of preparing phosphorus sulphides. By taking the right proportions of sulphur and phosphorus, the compounds P_4S_3 , P_4S_7 , P_4S_{10} , and P_2S_5 can be made in this manner. No traces of a disulphide or trisulphide P_3S_6 have been found in the freezing-point curves of phosphorus and sulphur; and the substance previously regarded as P_3S_6 (or PS_2) is probably a mixture. Several other sulphides have been reported, but some are in all probability mixtures.

Phosphorus pentasulphide, P_2S_5 .—The crude sulphide made by Kekulé's process is distilled in a current of carbon dioxide. A pale yellow crystalline mass is obtained which melts between 275° and 276° , and boils at 530° . The vapour density 7.67 (air = 1) corresponds with $7.67 \times 28.9 = 221.7$ (oxygen = 32). The theoretical value for P_2S_5 is 224 ($O_2 = 32$). This also agrees with the raising of the boiling point of solutions of the pentasulphide in carbon disulphide. This sulphide is hydrolyzed by water: $P_2S_5 + 8H_2O = 2PO(OH)_3 + 5H_2S$. It is used for replacing the oxygen in many hydroxyl compounds by sulphur. Thus with water, as we have just seen, $H-OH$ forms $H-SH$; and with alcohol, C_2H_5OH , it forms ethyl mercaptan, C_2H_5SH . With phosphorus pentachloride, PCl_5 , it forms thiophosphoryl chloride, $PSCl_3$, a colourless liquid boiling at 125° . This is the sulphur analogue of phosphoryl chloride. Phosphorus pentasulphide reacts with ammonia forming phosphorus hexamine pentasulphide, $P_2S_5 \cdot 6NH_3$, which, when gradually heated to redness in hydrogen or ammonia gas, furnishes a white odourless solid, phosphorus nitride, P_3N_5 . If the temperature has reached 850° , the colour varies from pure white to dark red. It is chemically inactive at ordinary temperatures, but at high temperatures it acts as a powerful reducing agent in consequence of its decomposition into phosphorus and nitrogen. *In vacuo*, this decomposition begins at 760° . Phosphorus nitride is hydrolysed by hot water forming phosphoric acid and ammonia.

Phosphorus sesquisulphide, tetraphosphorus trisulphide, P_4S_3 .—The crude sulphide prepared by Kekulé's process is a grey crystalline mass which can be purified by crystallization from solution in carbon disulphide or phosphorus trichloride, PCl_3 , in the form of rhombic prisms; or the crude product can be purified by distillation *in vacuo*. The vapour density of the sulphide corresponds with P_4S_3 . It dissolves in alkaline sulphides, and is slowly attacked by water. At 100° it inflames in air. It melts at 172.5° to a reddish liquid which boils at 407° or 408° . It is used in making "non-poisonous" safety matches. The freezing-point curve of mixtures of P_4S_3 and P_2S_5 shows a well-defined maximum corresponding with phosphorus heptasulphide, P_4S_7 . The same substance is formed when a carbon disulphide solution of the two sulphides is warmed above 100° . When the heptasulphide is heated to boiling with sulphur and naphthalene, phosphorus decasulphide, P_4S_{10} , is formed.

§ 7. Matches.

Common friction matches.—These are made by cutting soft wood into the required shape by machinery. One end of the strip is dipped into some inflammable substance—paraffin or sulphur, and then into a paste made from yellow phosphorus, manganese dioxide, glue, and colouring matter. The first friction matches were made with potassium chlorate and yellow phosphorus, and, in 1837, Preshal substituted lead dioxide, and later, a mixture of red-lead and manganese dioxide as safer oxidizing agents—potassium chlorate, nitre, etc. The matches are then dried. The glue protects the phosphorus from oxidation, but by rubbing the head of the match on a rough surface, sufficient heat is generated to ignite the phosphorus in contact with the oxidizing agent. The burning phosphorus

ignites the sulphur or paraffin, and this in turn fires the wood. Cotton threads dipped in paraffin are used in place of wood to form the so-called *wax vestas*.

Safety matches.—In spite of the greatest care, the phosphorus disease used to prevail in match factories using yellow phosphorus. The substitution of red for yellow phosphorus is far less dangerous to the health of the worker. About 1848, Böttger made the so-called safety matches with a "head" free from phosphorus. The head of the match is then made from a coloured mixture of antimony sulphide, potassium chlorate, and glue. Other oxidizing agents are used—red lead, potassium bichromate, etc. The surface on which the match is to be rubbed for ignition is covered with a mixture of red phosphorus, powdered glass, and glue. These matches—called safety matches—have the disadvantage that they can only be ignited by friction on a prepared surface or if quickly rubbed on a smooth non-conducting surface like glass or slate; but they are not liable to ignite by accidental friction.

Safety friction matches.—Matches can be made which strike on any surface by using Schenk's scarlet phosphorus, or phosphorus sesquisulphide— P_4S_3 —in place of ordinary phosphorus. These matches are not so liable to accidental ignition as ordinary friction matches, and they are made without risk of phosphorus poisoning.

Questions.

1. It is found that in 11.16 litres (calculated at 0°C . and 760 mm.) of any gaseous compound of phosphorus, there is never less than 15.5 grams of phosphorus present. Also that this volume of the vapour of phosphorus itself, under the same conditions, weighs 62 grams. State what conclusions may be drawn from the above data, with reference to the atomic and molecular weight of phosphorus.—*Cambridge Senior Locals*.

2. What volume changes occur where (i) Sulphur is heated in nitrous oxide; (ii) Phosphorus is heated in nitric oxide; (iii) Sodium is heated in gaseous hydrogen chloride; (iv) Potassium is heated in gaseous ammonia? Give equations. What inferences can be drawn from the observed facts as to the formulæ of the bodies formed and destroyed?—*Owens Coll.*

3. Describe and explain the changes which take place when (a) sulphur, (b) phosphorus, (c) iodine, (d) zinc, are severally boiled with concentrated potassium hydroxide solution.—*London Univ.*

4. Describe how from a specimen of calcium phosphate you would propose to make crystalline preparations of (a) calcium chloride, (b) ammonium magnesium phosphate. Give equations for the reactions involved.—*London Univ.*

5. What are the formula and name of the salt having the following percentage composition? Calcium, 38.72; phosphorus, 20.0; oxygen, 41.28.—*Glasgow Univ.*

6. Define the term "acid." In the light of your definition give reasons for assigning or refusing the name "acid" to aqueous solutions of the following: hydrogen sulphide, ordinary sodium phosphate, sodium bisulphate, copper sulphate, and alcohol ($\text{C}_2\text{H}_5\text{OH}$).—*Board of Educ.*

7. Describe the chlorides of phosphorus and give all the details essential for the preparation of a pure specimen of phosphorus pentachloride.—*Board of Educ.*

8. Describe how the law of mass action can be applied to the dissociation of phosphorus pentachloride, both by itself, and in the presence of chlorine gas, and state and explain what influence change of pressure and temperature have on this dissociation.—*Madras Univ.*

9. How do you account for the ignition of an ordinary lucifer match and of a safety match?—*London Univ. Metric.*

CHAPTER XXXI

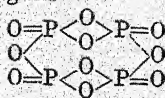
THE OXIDES AND ACIDS OF PHOSPHORUS

§ 1. Phosphorus Pentoxide.

WHEN phosphorus is burnt in an excess of dried air or oxygen, white clouds of phosphorus pentoxide condense as a voluminous powder—*flowers of phosphorus*. Small quantities of other oxides of phosphorus may be formed at the same time, and these impart a slight garlic-like smell to the product. Pure phosphorus pentoxide has no smell. Traces of the lower oxides of phosphorus present in commercial phosphorus pentoxide can be oxidized by re-subliming the powder in a current of dry oxygen so that the mixture passes over warm platinized asbestos.

Phosphorus pentoxide sublimes very slowly at 50° , but at 250° the sublimation is rapid. The vapour density at 1400° , by V. Meyer's process, corresponds with the molecular weight P_4O_{10} , but the simpler formula, P_2O_5 , is generally employed for the sake of convenience; and in virtue of our ignorance of the molecular formula of the solid oxide. Phosphorus pentoxide can be melted by heating it quickly. There are said to be three varieties: crystalline, amorphous, and vitreous.

Phosphorus pentoxide is extremely hygroscopic, and it absorbs moisture from the air very quickly; hence its use for drying gases. When thrown into water, combination occurs with a hissing sound resembling the quenching of red-hot iron, and metaphosphoric acid is formed: $P_2O_5 + H_2O = 2HPO_3$. When this solution is boiled with water, or when the phosphorus pentoxide is thrown into hot water, orthophosphoric acid is formed: $P_2O_5 + 3H_2O = 2H_3PO_4$. The heat of solution of the crystalline modification of phosphorus pentoxide is 41 Cals.; of the amorphous powder, 34 Cals.; of the vitreous variety, 29 Cals.; and the product of combustion of phosphorus with oxygen has a heat of solution of 34 Cals. The affinity of phosphorus pentoxide for moisture is so great that it can withdraw the elements of water from many organic and inorganic substances, *e.g.* it converts nitric acid into nitrogen pentoxide, etc. It is generally taken to resemble nitrogen pentoxide in constitution: $O_2 \equiv P - O - P \equiv O_2$, but this cannot apply to the molecule P_4O_{10} . If P_4O_{10} be the correct molecular formula of this oxide, the graphic formula (phosphorus quinquevalent) will be that shown in the diagram



Phosphoric oxide, P_4O_{10}

The corresponding acid— $\text{P}(\text{OH})_5$ —is not known; although the equivalent chloride, PCl_5 , is known.

§ 2. Orthophosphoric Acid and the Orthophosphates.

Orthophosphoric acid, H_3PO_4 , is formed when phosphorus pentoxide is dissolved in water, and the solution is boiled. It is also formed when one part of red phosphorus is boiled with 16 parts of nitric acid, specific gravity between 1.20¹ and 1.25, in a flask fitted with a reflux condenser, and a ground glass joint at the neck, as illustrated in Fig. 241, so that the nitric acid which is volatilized may be returned to the flask. When the phosphorus is all oxidized, the solution is evaporated to dryness, and the residue is finally heated in a platinum dish to a temperature not exceeding 180° to make sure that all the nitric acid is driven off. Orthophosphoric acid is also made by the action of sulphuric acid on bone ash as described for the manufacture of phosphorus.

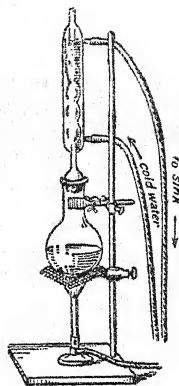
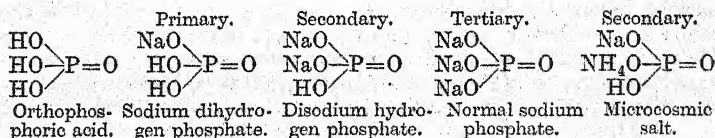


Fig. 241.—Preparation of Orthophosphoric Acid.

Properties of orthophosphoric acid.—By concentration *in vacuo*, or by heating to 140°, the acid can be obtained in six-sided prismatic crystals belonging to the rhombic system. The acid melts at 42.3°, and dissolves readily in water. It is tribasic— H_3PO_4 —and it forms three series of salts—normal or tertiary, secondary, and primary according as all, two, or one of its hydrogen atoms are replaced by an equivalent radicle:



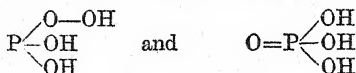
The normal salt is alkaline to litmus, the secondary salt is almost neutral, and the primary salt is acid. The hydrogen atom may be replaced by different radicles. Thus the secondary acid salt—ammonium sodium hydrogen phosphate, also called microcosmic salt—is illustrated by the graphic formula above. Two hydrates have been reported: $2\text{H}_3\text{PO}_4 \cdot \text{H}_2\text{O}$ and $10\text{H}_3\text{PO}_4 \cdot \text{H}_2\text{O}$. Orthophosphoric acid is comparatively weak, and appears to furnish the ions $\text{H}_3\text{PO}_4 \rightleftharpoons \text{H}^+ + \text{H}_2\text{PO}_4^- \rightleftharpoons 2\text{H}^+ + \text{HPO}_4^{2-}$ at moderate dilutions. Even at extreme dilutions the acid is only partially resolved into $3\text{H}^+ + \text{PO}_4^{3-}$. The acid functions of the three hydrogen ions vary greatly. The first ion has the strongest acid functions, the third, the least. To still further emphasize the idea developed in our study of neutralization, p. 170, the behaviour of the three sodium orthophosphates towards their indicators is given in Table XLIV.

¹ A stronger acid may cause an explosion, and with a weaker acid the action is slow. A trace of iodine will accelerate the action. The nitric acid should be free from sulphuric acid.

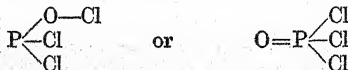
TABLE XLIV.—BEHAVIOUR OF THE SODIUM ORTHOPHOSPHATES TO INDICATORS.

Sodium orthophosphate.	Methyl orange.	Phenolphthalein.	Litmus.
Na_3PO_4 Na_2HPO_4 NaH_2PO_4	alkaline alkaline neutral	alkaline neutral acid	alkaline alkaline acid

The constitution of phosphoric acid.—Aqueous solutions of phosphoric acid contain the polymerized molecule $(\text{H}_3\text{PO}_4)_2$. The structural formula for phosphoric acid will depend upon what view is adopted about the valency of phosphorus—ter- or quinquevalent. In the former case, we have



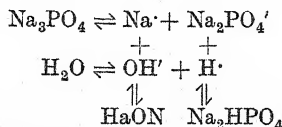
in the latter. The oxychloride of phosphorus must be either



Phosphorus is undoubtedly quinquevalent in the pentafluoride and also in the pentachloride. Accordingly, the latter formula is usually considered, without direct proof, to be the more probable, and hence it has been employed in what precedes. Phosphoric acid is formed by treating the oxychloride, $\text{O}=\text{PCl}_3$, with water, and therefore $\text{O}=\text{P}(\text{OH})_3$ is the most probable structural formula for phosphoric acid. With the notation discussed under periodic acid, "orthophosphoric acid" should be written $\text{P}(\text{OH})_5$, but this acid and its salts are unknown. The first dehydration product of $\text{P}(\text{OH})_5$, stable at ordinary temperatures, is called orthophosphoric acid $\text{O}=\text{P}(\text{OH})_3$; the corresponding chloride is POCl_3 , phosphoryl chloride. Just as the substitution of an alkyl radicle R' in place of hydroxyl in sulphuric acid furnishes a series of sulphonic acids, so here, the replacement of an alkyl radicle in phosphoric acid furnishes a series of phosphinic acids, $\text{R}.\text{PO}(\text{OH})_2$. This is a convenient place to emphasize the fact that compounds with the group $\text{O}=\text{P}\equiv$ are more stable than compounds with the group $\text{Cl}_2=\text{P}\equiv$. Phosphorus pentachloride, $\text{Cl}_2=\text{PCl}_3$, for instance, is readily dissociated by heat under conditions where $\text{O}=\text{PCl}_3$ is stable.

Sodium orthophosphates.—The disodium phosphate is the ordinary sodium phosphate used in the laboratory in testing for magnesium salts. It is obtained in the form of monoclinic prisms with 12 molecules of water of crystallization— $\text{Na}_2\text{HPO}_4.12\text{H}_2\text{O}$ —by adding sodium carbonate to phosphoric acid until the solution is just alkaline, and evaporating the solution until it deposits crystals. The crystals effloresce in air, and melt at 35° . The dodecahydrate passes into the heptahydrate at about 35° ; into the dihydrate at about 48° ; and into the anhydrous salt at about 95° . At 0° , 100 grams of water dissolve 2.5 grams of the salt; 82 grams, at 50° ; and 99, at 100° . If a mixed solution of sodium hydroxide and disodium phosphate be evaporated until the liquid crystallizes on cooling,

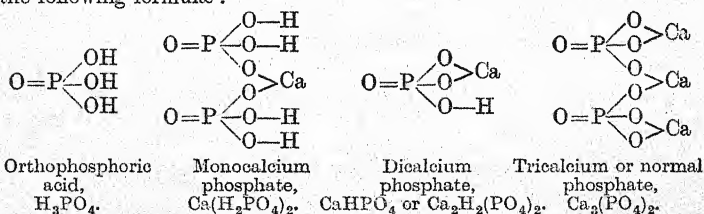
hexagonal or rhombic crystals of the normal sodium phosphate separate with 12 molecules of water of crystallization— $\text{Na}_3\text{PO}_4 \cdot 12\text{H}_2\text{O}$. Several other hydrates of this salt are known. The aqueous solution of the normal phosphate is alkaline—probably owing to the reaction: $\text{H}_2\text{O} + \text{Na}_3\text{PO}_4 \rightleftharpoons \text{Na}_2\text{HPO}_4 + \text{NaOH}$. In the language of ions, this is represented:



If phosphoric acid be added to the disodium phosphate until the liquid gives no precipitate with barium chloride, the solution when evaporated gives crystals of sodium dihydrogen phosphate— $\text{H}_2\text{NaPO}_4 \cdot \text{H}_2\text{O}$. The crystals are dimorphous; both forms belong to the rhombic system. Aqueous solutions of this salt are acid. If a mixture of, say, 6 grams of ammonium chloride, and 36 grams of the disodium phosphate be dissolved in as little hot water as possible, and the solutions be mixed, crystals of sodium ammonium hydrogen phosphate, microcosmic salt— $\text{HNaNH}_4\text{PO}_4 \cdot 4\text{H}_2\text{O}$ —will separate from the solution on cooling.

Silver orthophosphate, Ag_3PO_4 .—By mixing solutions of silver nitrate and sodium orthophosphate, a yellow precipitate of silver phosphate is obtained. The precipitate is soluble in nitric acid, and since nitric acid is produced during the precipitation of silver phosphate: $3\text{AgNO}_3 + \text{Na}_2\text{HPO}_4 = \text{Ag}_3\text{PO}_4 + 2\text{NaNO}_3 + \text{HNO}_3$, the precipitation will be incomplete. Silver phosphate also dissolves in aqueous ammonia.

Calcium orthophosphates.—Calcium phosphate is one of the most important salts of phosphoric acid. Pure crystalline tricalcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$, has not yet been found in nature, or prepared in the laboratory. The nearest approach to pure calcium phosphate is made by adding sodium phosphate to a solution of calcium chloride in the presence of ammonia. Calcined bones contain the equivalent of from 60 to about 80 per cent. of the normal phosphate. Several more or less impure calcium phosphates occur in nature, see "the occurrence of phosphorus." The normal and the two acid phosphates are related to orthophosphoric acid as indicated in the following formulæ:



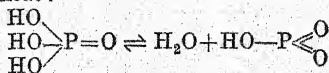
Normal calcium phosphate is very sparingly soluble in water: 100 parts of water dissolve about 0.003 part of the solid. The presence of alkalis decreases the solubility; while the presence of neutral salts—like sodium chloride, sodium nitrate—and acids—like carbon dioxide—increase its solubility in water. The dissolution of normal calcium phosphate in water—particularly in boiling water—is really a kind of hydrolysis, for a basic

Boiling water transforms the pyrophosphoric acid back into the orthophosphoric acid. Pyrophosphoric acid is tetrabasic. The ethyl salt, $(C_2H_5)_4P_2O_7$, has a molecular weight corresponding with the tetrabasic acid when its effect on the boiling point of chloroform is measured. Sodium pyrophosphate is formed when the monohydrogen salt is calcined: $2Na_2HPO_4 = H_2O + Na_4P_2O_7$. Magnesium ammonium phosphate is readily transformed into magnesium pyrophosphate by calcination, as indicated above: $2(NH_4)MgPO_4 = Mg_2P_2O_7 + H_2O + 2NH_3$. The pyrophosphates are fairly stable salts; they pass into the orthophosphates when boiled with sulphuric acid. Curiously enough, tetrabasic pyrophosphoric acid only forms quaternary or normal pyrophosphates, say $Na_4P_2O_7$, etc.; and the secondary pyrophosphates, say $Na_2H_2P_2O_7$, etc. The primary and tertiary salts are unknown. Pyrophosphoryl chloride, $P_2O_3Cl_4$, corresponds with pyrophosphoric acid, $P_2O_3(OH)_4$. The molecular weight of pyrophosphoric acid, prepared by the action of hydrogen sulphide on lead pyrophosphate, is $H_4P_2O_7$; that prepared by dehydrating orthophosphoric acid has a molecular weight corresponding with a mixture of $(H_4P_2O_7)_4$ and $(H_4P_2O_7)_5$.

Magnesium ammonium phosphate, $Mg(NH_4)PO_4 \cdot 6H_2O$.—This salt is obtained as a crystalline precipitate when a solution of a soluble phosphate is mixed with ammonia, ammonium chloride, and magnesium sulphate or chloride; and conversely, if ammonium phosphate be added to an ammoniacal solution of magnesium, the same salt is precipitated. These reactions are accordingly used for the determination of phosphorus, and of magnesium. When the precipitated magnesium ammonium phosphate is filtered, washed, and calcined, it is converted into magnesium pyrophosphate— $Mg_2P_2O_7$. The weight of the pyrophosphate enables the amount of phosphorus or of magnesium to be computed. Since the alkaline earths would also be precipitated as phosphates with the magnesium phosphate, if they be present, the analyst removes the alkaline earths—say by precipitation as carbonates in the presence of ammonium chloride in which magnesium carbonate is soluble—before the magnesium phosphate is precipitated. The same remark also applies to other precipitants for magnesium; they also precipitate the alkaline earths.

EXAMPLE.—How much magnesia, MgO , and how much phosphoric oxide, P_2O_3 , would 0.5 grm. of magnesium pyrophosphate represent? Ansr. 0.181 grm. of MgO , and 0.319 grm. P_2O_3 .

Metaphosphoric acid, HPO_3 .—This acid is formed as a viscous solid when phosphorus pentoxide is left exposed to moist air; and it is also obtained by heating ammonium orthophosphate, or pyro- or orthophosphoric acid to a red heat:



Metaphosphoric acid apparently bears the same relation to phosphorus that nitric acid bears to nitrogen. Metaphosphoric acid is a transparent vitreous solid, sometimes called "glacial phosphoric acid." It fuses at about $38^\circ\text{--}41^\circ$, and the commercial acid is usually cast in sticks. It is readily soluble in water, and the solution slowly passes into the orthoacid; the change proceeds rapidly on boiling the solution. Several polymers— HPO_3 , $(HPO_3)_2$, and $(HPO_3)_3$ —have been obtained.

Vapour density determinations of the metaphosphoric acid prepared as indicated above, correspond with the formula $\text{H}_3\text{P}_2\text{O}_6$, and in consequence, the acid may be a dimetaphosphoric acid. A variety prepared by the action of hydrogen sulphide on the lead salt has the molecular composition HPO_3 .

Consonant with the prefixes adopted for the periodic acids, $\text{P}(\text{OH})_5$ should be called orthophosphoric acid; $\text{PO}(\text{OH})_3$ mesophosphoric acid; and $\text{PO}_2(\text{OH})$ metaphosphoric acid. Pyrophosphoric acid would then be dimesophosphoric acid. The reason this acid is called "pyro-" acid will be obvious from its mode of formation—Greek $\pi\upsilon\rho$ (pyr), fire.

The metaphosphates.—Sodium metaphosphate, NaPO_3 , is formed by igniting either dihydrogen sodium phosphate, or hydrogen sodium ammonium phosphate, or dihydrogen sodium pyrophosphate. Metaphosphoric acid is itself monobasic, and it forms a series of polybasic salts which may be regarded as derivatives of the hypothetical polymerized metaphosphoric acids. For instance:

ACIDS.		SALTS.	
HPO_3 , Monometaphosphoric acid		NaPO_3 , Sodium metaphosphate	
$(\text{HPO}_3)_2$, Dimetaphosphoric acid		$\text{K}_2\text{P}_2\text{O}_6$, Potassium dimetaphosphate	
$(\text{HPO}_3)_3$, Trimetaphosphoric acid		$\text{Na}_3\text{P}_3\text{O}_9$, Sodium trimetaphosphate	
$(\text{HPO}_3)_4$, Tetrametaphosphoric acid		$\text{Pb}_2\text{P}_4\text{O}_{12}$, Lead tetrametaphosphate	
$(\text{HPO}_3)_5$, Pentametaphosphoric acid		$(\text{NH}_4)_5\text{P}_5\text{O}_{15}$, Amm. pentametaphosphate	
$(\text{HPO}_3)_6$, Hexametaphosphoric acid		$\text{Na}_6\text{P}_6\text{O}_{18}$, Sodium hexametaphosphate	
...

There is some reason to doubt if many of the polyphosphates are anything more than mixtures of pyro- and meta-phosphates. Ferrous metaphosphate, $\text{Fe}(\text{PO}_3)_2$, has been obtained by the action of fused metaphosphoric acid on iron, or on ferrous salts in an atmosphere of carbon dioxide. It is a white powder with a greenish tinge, insoluble in nitric or hydrochloric acid, but oxidized by hot sulphuric acid. Molybdenum metaphosphate was prepared in a similar way, but the process did not furnish chromous metaphosphate, but chromic metaphosphate, $\text{Cr}(\text{PO}_3)_3$, was always obtained.

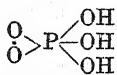
Distinguishing tests for the phosphoric acids.—The three phosphoric acids and their salts are distinguished by the difference in their behaviour towards silver nitrate, barium nitrate, albumen, etc.:

TABLE XLV.—REACTIONS OF THE PHOSPHORIC ACIDS.

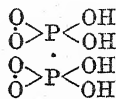
	Orthophosphoric acid.	Pyrophosphoric acid.	Metaphosphoric acid.
Silver nitrate . .	Canary-yellow pp.	White cryst. pp.	White gelatinous.
Barium salts . .	No pp. (if alkaline, white pp.).	No pp. (if alkaline, white pp.).	White precipitate.
Albumen . . .	Nil.	Nil.	Coagulated.
Zinc acetate . .	Nil.	Insoluble pp.	Nil.
Aluminium salts } Chromium salts }	Soluble in acetic acid.	Insoluble in acetic acid.	Insoluble in acetic acid.
Cobalt salts . .	Blue; soluble in acetic acid.	Red; insoluble in acetic acid.	Red; insoluble in acetic acid.
Copper salts . .	Nil.	Precipitate.	Nil.
Bismuth salt alkaline (solution) .	Nil.	Nil.	Precipitate.

All the phosphoric acids give a canary-yellow precipitate with a large excess of a boiling solution of ammonium molybdate in nitric acid. The precipitate is soluble in phosphoric acid, and a large excess of the molybdate solution is needed or no precipitate may be formed. The same result is obtained with all the acids of phosphorus since the nitric acid oxidizes them to phosphoric acid. Arsenic oxyacids give a similar precipitate.

When phosphorus pentoxide is treated with 30 per cent. hydrogen peroxide, while cooled with ice-water, an acid—monoperphosphoric acid—resembling Caro's acid—possibly H_3PO_5 —is obtained; with pyrophosphoric acid, crystals with the empirical formula $\text{H}_4\text{P}_2\text{O}_8$, diperphosphoric acid, or simply perphosphoric acid corresponding with persulphuric acid, are formed. The probable graphic formulæ are not unlike those for persulphuric acids, excepting that sulphur is sexivalent, and phosphorus quinquivalent :



Monoperphosphoric acid,
 H_3PO_5



Diperphosphoric acid,
 $\text{H}_4\text{P}_2\text{O}_8$

History of the phosphoric acids.—R. Boyle discovered phosphoric acid in 1746, and A. S. Marggraf prepared it by dissolving *fleurs de phosphore* in water, and by heating phosphorus with nitric acid; he also described some chemical and physical properties. It was once thought that phosphoric acid existed in three isomeric forms: (1) the ordinary acid which gave a yellow precipitate with silver nitrate (A. S. Marggraf, 1746); (2) the product obtained by heating the ordinary acid which gave a white precipitate with silver nitrate (T. Clark, 1827); (3) the acid obtained by a thorough calcination of phosphoric acid and which gave a white precipitate with silver nitrate, and unlike the other two forms, coagulated a clear aqueous solution of albumen (J. J. Berzelius, 1816). Thomas Graham, 1833, proved that these three acids—which he called respectively ordinary, pyro- and metaphosphoric acids—were different modifications of phosphoric acid; and that they differed from one another by “the quantity of water combined with the acid.” Graham also found that when the acids were saturated with a base, three series of phosphates were obtained, one series contained one equivalent of the base per equivalent of the acid, a second series contained two equivalents of the base; and a third series contained three equivalents of the base.

Derivatives of orthophosphoric acid are known—e.g. the thiophosphates, and oxythiophosphates—where the oxygen atoms are wholly or partially replaced by sulphur atoms.

§ 4. Phosphorous Oxide or Phosphorus Trioxide and Phosphorous Acid.

Phosphorous oxide— P_4O_6 —mixed with a large excess of phosphorus pentoxide, is formed when phosphorus is burnt in a limited supply of air. Some red phosphorus and possibly a lower oxide of phosphorus are formed at the same time. For this experiment, the phosphorus is placed in a

glass tube, *A*, which is bent as shown in Fig. 242, and fitted into one end of a long tube cooled by a jacket, *B*, containing water at 60°. The cooled tube is fitted to a U-tube, *C*, immersed in a freezing mixture; a plug of glass-wool is placed in the condenser tube near the U-tube. The phosphorus is ignited, and a slow stream of air is drawn through the apparatus by means of an aspirator connected to the U-tube. The

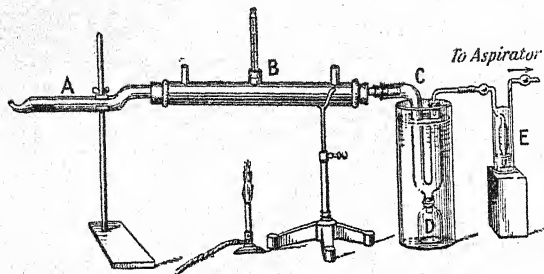


FIG. 242.—The Preparation of Phosphorus Trioxide— P_4O_6 .

phosphorus pentoxide is arrested by the glass-wool, and phosphorous oxide passes into the U-tube, where it is condensed into a white crystalline mass. The wash-bottle, *E*, with concentrated sulphuric acid, protects the product from moisture. At the end of the experiment, the solid in the U-tube can be melted and run into the bottle *D*.

Properties.—Phosphorous oxide so formed is a mass of monoclinic prisms; it melts at 22·5° to a waxy solid, and boils at 173°. The vapour density of the solid corresponds with the molecule P_4O_6 , although the simpler formula, P_2O_3 , is often used. The lowering of the freezing points of solutions of phosphorous oxide in benzene corresponds with a molecular weight of 227; and solutions in naphthalene, with a molecular weight 218. The theoretical value for P_4O_6 is 220; the graphic formula being

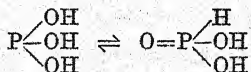


where the phosphorus is assumed to be tervalent. Phosphorous oxide smells like garlic and it is poisonous. When exposed to the air it is gradually oxidized to the pentoxide, and when placed in warm oxygen it bursts into flame. It also ignites spontaneously in chlorine. It is slowly attacked by cold water, forming phosphorous acid, H_3PO_3 , and with hot water it forms red phosphorus, a lower oxide of phosphorus, phosphoric acid, and phosphine. When heated with sulphur to about 159° in an atmosphere of carbon dioxide or nitrogen, an additional product is formed which resembles phosphorus pentoxide with two oxygen atoms replaced by sulphur atoms. It is therefore phosphorus trioxydisulphide, $P_2O_3S_2$, or rather, since the vapour density is 343·8 ($H_2 = 2$), $P_4O_6S_4$. The compound is rapidly decomposed by water forming hydrogen sulphide and metaphosphoric acid, which, in turn, reacts with more water forming orthophosphoric acid: $P_4O_6S_4 + 6H_2O = 4HPO_3 + 4H_2S$.

Phosphorous acid, H_3PO_3 .—This acid is formed by the action of cold water upon phosphorous oxide: $P_4O_6 + 6H_2O = 4H_3PO_3$; by the action of water on phosphorus trichloride: $PCl_3 + 3H_2O = 3HCl +$

H_3PO_3 ; or by passing a stream of chlorine through water beneath which phosphorus is melted. The phosphorus trichloride formed in the latter reaction is at once decomposed by the water into phosphorous acid. The solution is evaporated until the temperature has reached 180° , when, on cooling, it solidifies to a crystalline solid which melts at 70.1° . When heated, phosphorous acid decomposes into phosphine and ortho-phosphoric acid: $4\text{H}_3\text{PO}_3 = 3\text{H}_2\text{PO}_4 + \text{PH}_3$. Phosphorous acid is a powerful reducing agent. It reduces silver nitrate to metallic silver; gold chloride to metallic gold; and copper sulphate to *metallic copper*. It also absorbs oxygen, forming phosphoric acid. Phosphorous acid is reduced by zinc and hydrochloric acid to phosphine: $\text{H}_3\text{PO}_3 + 3\text{Zn} + 6\text{HCl} = 3\text{ZnCl}_2 + 3\text{H}_2\text{O} + \text{PH}_3$. The zinc and hydrochloric acid of course furnish nascent hydrogen.

Constitution.—Phosphorous acid usually behaves as a dibasic acid, but a salt—sodium phosphite, Na_3PO_3 —corresponding with a tribasic acid has been reported. Consequently, it is not possible to say definitely whether the formula of phosphorous acid should be written $\text{P}(\text{OH})_3$, or $\text{OPH}(\text{OH})_2$, since the evidence from different sources is contradictory. The formation of the acid from phosphorus trichloride points to the formula $\text{P}(\text{OH})_3$, and the dibasicity of the acid to the formula, $\text{O} : \text{PH}(\text{OH})_2$. The latter formula appears the more probable, and under special conditions the hydrogen directly attached to the phosphorus atom may have acidic properties, in the same way that the hydrogen of ammonia can be replaced by other radicles. Triethyl phosphite, $\text{P}(\text{OC}_2\text{H}_5)_3$, boiling between 155° and 156° (760 mm.), appears to be derived from a symmetrical tribasic phosphorous acid; though the scarcity of salts of the tribasic acid, and the existence of the unsymmetrical $\text{O} : \text{PC}_2\text{H}_5(\text{OC}_2\text{O}_5)_2$ show that two of the replaceable hydrogen atoms probably differ from the third. As in the case of sulphurous and nitrous acids, the facts are explained by assuming the desmotropic change:



As indicated previously, phosphorus may be both ter- and quinequevalent; but it is generally assumed that the properties of phosphorous acid correspond best with a quinequevalent phosphorus, and that the formula is $\text{O} : \text{PH}(\text{OH})_2$. The conclusion that phosphorous acid is dibasic is in agreement with the heat of neutralization of the acid with sodium hydroxide. J. Thomsen found that

Gram-molecules of base	1	2	3
Heat evolved in Cals.	14.8	27.1	34.0

Increasing the proportion of alkali beyond this stage has no marked effect on the thermal value of the reaction because no further combination occurs. The results with pyrophosphoric, sulphuric, carbonic, boric, and other acids agree with the basicities assigned to them in this book. Hence, if the molecular weight of an acid be known, its basicity can be deduced from the heats evolved when the acid is treated with increasing proportions of alkali. With metaphosphoric acid the results appear to be exceptional, for they indicate a higher basicity than corresponds with HPO_3 , this is because the

meta-acid is unstable in aqueous solution, and is continuously passing into the ortho-acid. With phosphorous acid,

Gram-molecules of base	1 ¹	2	3
Heat evolved in Cals.	14.8	28.5	28.8

The very slight effect of the third gram-molecule of sodium hydroxide agrees with the assumption that the acid is dibasic.

The phosphites.—The salts of phosphorous acid—the phosphites—are soluble in water, and those of the alkali metals have an acid reaction. Solutions of phosphites give a precipitate with baryta and with lime water; they also precipitate metals from salts of silver, gold, and mercury.

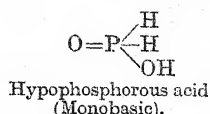
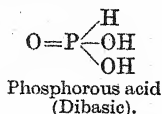
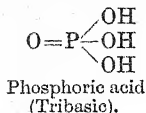
Polyphosphorous acids.—A number of polyphosphorous acids— $\text{H}_4\text{P}_2\text{O}_5$, $\text{H}_5\text{P}_3\text{O}_7$, $\text{H}_7\text{P}_5\text{O}_{11}$ —or rather salts corresponding with these acids, can be obtained by removing molecules of water from one or more molecules of the salts of phosphorous acid. These salts are analogous with the corresponding salts of phosphoric acid. Thus $\text{P}_2\text{O}_3 + 3\text{H}_2\text{O} = 2\text{H}_3\text{PO}_3$, or $\text{P}(\text{OH})_3$ —orthophosphorous acid; $\text{P}_2\text{O}_3 + 2\text{H}_2\text{O} = \text{H}_4\text{P}_2\text{O}_5$ —pyrophosphorous and $\text{P}_2\text{O}_3 + \text{H}_2\text{O} = 2\text{HPO}_2$ —metaphosphorous acid. The last-named acid is analogous with nitrous acid, HNO_2 . Pyrophosphorous acid, $\text{H}_4\text{P}_2\text{O}_5$, is formed by the action of water on phosphorus trichloride, PCl_3 , in the presence of hydrochloric acid: $2\text{PCl}_3 + 5\text{H}_2\text{O} = 6\text{HCl} + \text{H}_4\text{P}_2\text{O}_5$. The colourless crystalline mass melts at 38° and forms the ortho acid when treated with water. According to H. G. van de Stadt (1898), a metaphosphorous acid, HPO_2 , corresponding with one molecule of H_3PO_3 less one molecule of water, is formed in white feathery crystals when phosphine slowly oxidizes under reduced pressure. It is said to be converted into phosphorous acid by the action of water vapour; and to melt at a higher temperature than phosphorous acid. There is some doubt whether this is really metaphosphorous acid.

§ 5. Hypophosphorous Acid, and the Hypophosphites.

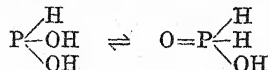
Preparation.—When phosphorus is boiled with a solution of barium hydroxide, barium hypophosphite, $\text{BaH}_2\text{P}_2\text{O}_4$, is formed: $3\text{Ba}(\text{OH})_2 + 8\text{P} + 6\text{H}_2\text{O} = 2\text{PH}_3 + 3\text{BaH}_2\text{P}_2\text{O}_4$. When this solution is treated with sulphuric acid, barium sulphate is precipitated. The clear aqueous solution is separated from the precipitate by filtration; and, when the solution is evaporated in vacuo at 80° – 90° it will deposit white crystals of the acid when cooled to 0° . The acid can be purified by recrystallization.

Properties.—The acid has the empirical formula $\text{H}_2\text{P}_2\text{O}_4$. The crystals melt at 26.4° . The acid is reduced by zinc and hydrochloric acid to phosphine. Hypophosphorous acid is a feeble monobasic acid which forms a series of salts called hypophosphites, where $\text{H}_2\text{P}_2\text{O}_4$ acts as a monad radicle, thus, sodium hypophosphite, $\text{NaH}_2\text{P}_2\text{O}_4$ (p. 716); barium hypophosphite, $\text{Ba}(\text{H}_2\text{P}_2\text{O}_4)_2$, etc. The acid and its salts act as energetic reducing agents, thus, with copper sulphate a red precipitate of copper hydride, Cu_2H_2 , is obtained. This shows that the reduction progresses a step further than is the case with phosphorous acid, and this reaction is one of the distinctive tests for hypophosphites. The acid and its salts also reduce salts of silver, gold, etc., to metals. The hypophosphites are oxidized to phosphates by oxidizing agents.

Constitution.—Assuming the quinquevalency of phosphorus, the relations between phosphoric, phosphorous, and hypophosphorous acids are well illustrated by the graphic formulæ :



The last formula emphasizes the monobasicity of the hypophosphorous acid, although it is probable that one of the hydrogen atoms is labile corresponding with the desmotropic change :



Inorganic derivatives of dibasic hypophosphorous acid are not known. The assumed monobasicity of this acid is confirmed by observations on the heat evolved when the acid is neutralized by sodium hydroxide. Thus, one gram-molecule of sodium hydroxide evolves 15.2 Cals. when mixed with a gram-molecule of hypophosphorous acid, while with two gram-molecules of sodium hydroxide, 15.3 Cals. are evolved. The tendency of phosphorus to pass into the $\text{O}=\text{P}\equiv$ form is illustrated by the action of heat on this acid, and on its salts, for phosphoric acid and phosphine are produced : $2\text{H}_3\text{PO}_2 = \text{OP}(\text{OH})_3 + \text{PH}_3$. Neither the anhydride nor the chloride corresponding with hypophosphorous acid is known.

§ 6. Phosphorus Tetroxide. Hypophosphoric Acid.

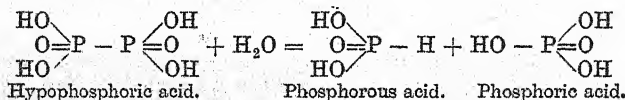
Phosphorus tetroxide, P_2O_4 .—When phosphorus trioxide, P_4O_6 , is heated in a sealed tube to about 440° , it decomposes into phosphorus tetroxide, P_2O_4 , which appears as a crystalline sublimate in the sealed tube. The tetroxide reacts with water, forming a mixture of phosphoric and phosphorous acids : $\text{P}_2\text{O}_4 + 3\text{H}_2\text{O} = \text{H}_3\text{PO}_3 + \text{H}_3\text{PO}_4$. Phosphorus tetroxide is therefore to be regarded as a mixed anhydride which bears the same relation to phosphorus than nitrogen tetroxide, N_2O_4 , bears to nitrogen. The latter, it will be remembered, forms nitrous and nitric acids with water.

Hypophosphoric acid, $\text{H}_4\text{P}_2\text{O}_6$, or H_2PO_3 .—When phosphorus is exposed to a limited supply of moist air, a mixed solution of phosphoric, phosphorous, and hypophosphoric acids is formed. The acidified liquid is neutralized with sodium hydroxide and a sparingly soluble sodium hypophosphate separates out. The acid is made by the electrolysis of water made slightly acid with sulphuric or formic acid, using copper phosphide as anode, and copper as cathode. A solution of this salt when treated with lead acetate gives an insoluble precipitate of lead hypophosphate, PbPO_3 . This salt is filtered from the solution, and washed with hot water. The precipitate is suspended in water, and a current of hydrogen sulphide passed through the solution. Lead sulphide, PbS , is precipitated, and a solution of free hypophosphoric acid is obtained. On evaporation, the excess of hydrogen sulphide is driven from the solution, but the evaporation

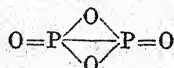
cannot be carried very far without decomposing the acid. Hence, the solution must be further evaporated in a desiccator *in vacuo* over sulphuric acid, Fig. 87. In time, tabular, rhombic crystals of the hydrate H_2PO_3 . H_2O separate. The crystals melt at 62° . Further desiccation of the crystals over sulphuric acid *in vacuo* gives the anhydrous acid, H_2PO_3 , melting at 70° . The acid is stable at ordinary temperatures, and is hydrolyzed by the mineral acids, forming a mixture of phosphoric and phosphorous acids. The acid is dibasic, and the aqueous solution, unlike phosphorous acid, does not possess reducing qualities.

Relation between hypophosphoric acid and phosphorus tetroxide.

—The fact that P_2O_4 furnishes phosphoric and phosphorous acids when treated with water is expressed by the structural formula $\text{O}=\text{P}\equiv\text{O}_3\equiv\text{P}$. If phosphorus tetroxide be the anhydride of hypophosphoric acid, we should expect: $\text{O}=\text{P}\equiv\text{O}_3\equiv\text{P} + 2\text{H}_2\text{O} \rightarrow (\text{OH})_2\text{PO}-\text{O}-(\text{OH})_2$. If the last formula really represents the structure of hypophosphoric acid, we should expect this acid to be a reducing agent like phosphorous acid, but it is not. Its decomposition into phosphoric and phosphorous acids is well represented by the following scheme:



Accordingly, phosphorus tetroxide, which also decomposes into the same two acids in contact with cold water, has probably the structural formula



With two molecules of water, this gives hypophosphoric acid, and with three molecules of water a mixture of phosphoric and phosphorous acids. Assuming that hypophosphoric acid (melting point 70°) corresponds with the above formula $(\text{OH})_2\text{OP}-\text{PO}(\text{OH})_2$; the hydrate (melting at 62°) is sometimes regarded as $(\text{HO})_4\text{P}-\text{P}(\text{OH})_4$. The molecular weight, deduced from the effect of the methyl esters, $(\text{CH}_3)_2\text{PO}_3$, on the boiling points of methyl bromide or iodide, corresponds with the formula H_2PO_3 , not $\text{H}_4\text{P}_2\text{O}_6$. The acid undergoes association in aqueous solutions. The freezing points of aqueous solutions of the acid and of the potassium salt agree with the doubled formula $\text{H}_4\text{P}_2\text{O}_6$; so also does the effect of dilution on the electrical conductivity of the sodium salts. There are, however, disturbances due to the decomposition of the salt in dilute solutions. The progressive neutralization of a solution of the acid with alkali hydroxide or ammonia favours the doubled formula. If the formula be H_2PO_3 , phosphorus seems to act as a quadrivalent element in $\text{O}=\text{P}=(\text{OH})_2$.

By exposing phosphorus pentachloride to the action of ammonia, and heating the product in air free from oxygen, or by heating phosphorus with ammonium chloride, a white powder called phospham, HPN_2 , is formed. The formula HPN_2 recalls that of hydrazoic acid, HN_3 , but the two compounds have not much in common. Phospham reacts with hot water forming metaphosphoric acid and ammonia: $\text{HPN}_2 + 3\text{H}_2\text{O} =$

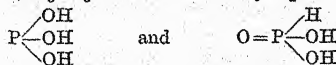
$\text{HPO}_3 + 2\text{NH}_3$. A curious set of six compounds is formed by heating ammonium chloride and phosphorus pentachloride in sealed tubes and subjecting the product to fractional distillation. All the compounds so separated have the empirical formula $(\text{PNCl}_2)_n$, where n may be 3, 4, 5, 6, 7, and some higher unknown value. These phosphonitrile chlorides yield a series of phosphimic acids when treated with water.

Summary.—To summarize, the more well-defined oxides and oxyacids of phosphorus are:

OXIDES.	ACIDS.
Phosphorus suboxide
Phosphorous oxide, P_4O_6 ...	Hypophosphorous acid, H_3PO_2 Orthophosphorous acid, H_3PO_3 Pyrophosphorous acid, $\text{H}_4\text{P}_2\text{O}_5$ Metaphosphorous acid, HPO_3
Phosphorus tetroxide, P_2O_4
Phosphorus pentoxide, P_2O_5 ...	Hypophosphoric acid, $\text{H}_4\text{P}_2\text{O}_6$ Orthophosphoric acid, H_3PO_4 Pyrophosphoric acid, $\text{H}_4\text{P}_2\text{O}_7$ Metaphosphoric acid, HPO_3 Monoperphosphoric acid, H_3PO_5 Diperphosphoric acid, $\text{H}_4\text{P}_2\text{O}_8$
...	...
...	...
...	...

Questions.

- How are phosphorus trioxide and phosphorus pentoxide respectively prepared? What happens when phosphorus pentoxide is dissolved in cold water and then the solution boiled? What is the action of heat on orthophosphoric acid?—*Aberdeen Univ.*
- Make the equation for the preparation of phosphine. Compare phosphine and ammonia. How may phosphorus be converted to phosphoric acid? Give equations.—*Univ. Pennsylvania, U.S.A.*
- What is the effect of heat on the following: (a) phosphorous acid, (b) potassium chlorate, (c) lead nitrate, (d) ammonium chloride?—*St. Andrews Univ.*
- Give some examples of constitutional formulae; and discuss the question of their use in chemical theory and investigation.—*New Zealand Univ.*
- How may the different modifications of phosphoric acid be obtained, and by what tests may they be distinguished?—*Aberdeen Univ.*
- Phosphorous acid, H_3PO_3 , has been variously represented as—



Discuss the question indicating the significance of the following facts: (1) That in most of its salts phosphorous acid is tribasic; (2) That an ether exists of the formula $\text{PO}_3(\text{C}_6\text{H}_5)_3$; (3) That phosphorous acid may be formed by the action of water upon phosphorus trichloride; (4) That phenyl phosphonitrile chloride, $\text{C}_6\text{H}_5\text{PCl}_2$, is formed by the action of the trichloride on benzene, C_6H_6 , and that when the body is treated with water, $\text{PO}_3\text{H}_2(\text{C}_6\text{H}_5)$ is produced which on treatment with PCl_5 affords the following reaction: $\text{OPH}(\text{OH})\text{C}_6\text{H}_5 + 2\text{PCl}_5 = \text{OPCl}_2(\text{C}_6\text{H}_5) + \text{POCl}_3 + \text{PCl}_3 + 2\text{HCl}$.—*London Univ.*

7. Give a brief account of the preparation, isolation, and chief characters of the several oxides of phosphorus; and compare together the known oxides of nitrogen, phosphorus, and arsenic.—*London Univ.*

8. Describe the preparation of phosphorus from bone ash; and give details, with a rough sketch of the apparatus you would employ, in order to convert this element into (a) phosphorus pentoxide, (b) phosphine, (c) phosphorus trichloride, (d) phosphonium iodide.—*London Univ.*

9. Give the names of the substances corresponding to the formulae in the following equation: $\text{P}_4 + 3\text{KOH} + 3\text{H}_2\text{O} = \text{PH}_3 + 3\text{KH}_2\text{PO}_2$; and calculate the volume of gas at 21° and 780 mm. pressure, and the weight of the salt which ought to be obtained by using 70 grams of phosphorus. Why do we write P_4 and not

$2P_2$? ($H = 1, O = 16, P = 31, K = 39$, one litre of hydrogen at 0° and 760 mm. pressure weighs 0.09 gram.)

10. Give an account of the preparation of orthophosphoric acid; show why it is considered to be a tribasic acid; and also by a structural formula how it is related to pyrophosphoric acid.—*London Univ.*

11. State "the law of mass action" and illustrate it by the following examples: (a) the action of water on antimony trichloride, (b) the action of steam on red hot iron.

12. How is yellow phosphorus converted into the red variety, and how can the reverse change be accomplished? What substances can be formed when phosphorus is (a) heated in air, (b) heated with nitric acid, and what is the relationship between them?—*Board of Educ.*

13. Describe briefly the preparation of sodium hypophosphite from phosphorus. Explain carefully the changes which occur when this salt is heated, and when solutions of mercuric chloride and acidified potassium permanganate, respectively, are mixed with its aqueous solution. What is the basicity of hypophosphorous acid and how has it been ascertained?—*Board of Educ.*

14. Give the formulæ of meta- and of common phosphoric acids. How could you distinguish these acids in aqueous solution? Illustrate by examples the chemical differences between the mono-, di-, and tri-acids.—*London Univ.*

15. Give a brief account of nitrogen and phosphorus, and their chief compounds, more especially with regard to their importance in animal and vegetable life.—*Sydney Univ.*

16. What is superphosphate? Describe its method of preparation on a commercial scale. How much superphosphate will be formed from 150 lbs. of pure bone ash, and how much would the P_2O_5 in a ton of superphosphate be worth at 4d. per lb.?—*Bombay Univ.*

17. What is meant by the equivalent of an element? In what sense can we use this term correctly with reference to a compound? Give the equivalents of iron, nitrogen, caustic soda, sodium sulphate, orthophosphoric acid.—*London Univ.*

CHAPTER XXXII

ARSENIC, ANTIMONY, AND BISMUTH

§ 1. Arsenic—Occurrence, Preparation, and Properties.

Atomic weight, As = 74.96; molecular weight, As_4 = 299.84. Ter- and quinquevalent. Melting point, under pressure, about 817° ; boiling point, 450° .

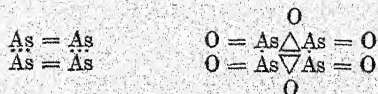
Occurrence.—The element occurs free, and combined in a great number of minerals—oxide, *arsenolite*, As_4O_6 ; sulphides, *realgar*, As_2S_2 ; *orpiment*, As_2S_3 ; *mispickel* or *arsenical pyrites*, FeAsS ; *cobaltite* or *cobalt glance*, CoAsS ; arsenides—*tin white cobalt*, CoAs_2 ; *arsenical iron*, FeAs_2 , and Fe_4As_3 ; *nickel glance*, NiAsS ; *kupfernickel*, NiAs . Arsenic also occurs in most samples of pyrites and hence it finds its way into sulphuric acid when sulphur dioxide is made by roasting pyrites. Arsenic is also found in commercial zinc; and in the smoke from coal when the coal contains pyrites; hence also arsenic finds its way into the atmosphere of towns, where it can be detected, particularly in foggy weather. Arsenic is also found in some mineral waters—e.g. Levico, Roncesgno, etc. A. Gautier says that traces also occur normally in the human body. The mere qualitative detection of arsenic does not therefore give much information unless the operation is more or less quantitative.

Preparation.—The element arsenic can be prepared by heating a mixture of the oxide, As_4O_6 , with powdered charcoal in a clay crucible: $\text{As}_4\text{O}_6 + 6\text{C} = 6\text{CO} + 4\text{As}$. The crucible is provided with a conical iron cap in which the arsenic sublimes. Most of the commercial arsenic is either a natural product, or else it is made by heating mispickel in a clay tube fitted half its length with an inner sheet-iron tube. The arsenic sublimes into the iron tube. By withdrawing and unrolling the tube, the element arsenic is obtained: $\text{FeAsS} = \text{FeS} + \text{As}$. The arsenic so obtained is not very pure. It is purified by resublimation from a mixture of the crude element and charcoal.

Properties.—Ordinary arsenic is a steel-grey metallic-looking substance, which forms hexagonal rhombohedral crystals with a bright lustre. It is called grey arsenic or γ -arsenic to distinguish it from two other allotropic modifications. Grey arsenic is brittle, and, like the metals, it is a good conductor of heat. Its specific gravity—5.727—is higher than typical non-metals. In general physical properties grey arsenic resembles the metals, but otherwise it is classed with phosphorus among the non-metals. At atmospheric pressures it sublimes very slowly at about 100° , and very rapidly at a dull red heat, without melting. If heated under pressure, it melts about 817° , but under ordinary pressures it sublimes without melting. The vapour is lemon-yellow, and it smells like garlic.

Arsenic is not altered by exposure to dry air but in moist air a surface film of oxide is formed. At 180° it burns with a bluish flame, forming arsenious oxide, As_2O_3 . It is readily oxidized by concentrated nitric and sulphuric acids. Dilute sulphuric acid has very little action on arsenic, while hot concentrated sulphuric acid dissolves arsenic forming sulphur dioxide and *probably* a very unstable arsenic sulphate, $As_2(SO_4)_3$, which immediately decomposes into the oxide. Dilute nitric acid in the cold has very little action, but the hot acid oxidizes the element to arsenic acid— H_3AsO_4 , or As_2O_5 . The reactions can be represented: $6As + 10HNO_3 = 3As_2O_5 + 5H_2O + 10NO$. Concentrated nitric acid and *aqua regia* also form arsenic acid. Arsenic is not appreciably attacked by hydrochloric acid in the absence of air; but if air be present, it is slightly soluble. As_2O_3 is said to be first formed and then dissolved as arsenic trichloride, $AsCl_3$. The element combines directly with chlorine at ordinary temperatures, forming arsenic trichloride. If a piece of clean copper foil be boiled with an arsenic salt, and hydrochloric acid, a grey film of arsenic is deposited on the copper. Several other metals, if present, give a similar film, but by heating the dried foil in a narrow glass tube, the arsenic is oxidized and sublimed on to the cooler parts of the tube in the form of octahedral crystals of arsenic trioxide—**Reinsch's test**. Arsenic unites with almost all the metals, forming arsenides—*e.g.* iron— $FeAs_2$; cobalt— $CoAs_2$; nickel— $NiAs$, etc. Arsenides of the metals can be made by heating finely powdered metal and arsenic in a sealed tube. Manganese arsenide, $MnAs$, is made in this way; the excess of arsenic is removed with chlorine or bromine, and the excess of manganese by hydrochloric acid. It is a greyish-black powder, strongly magnetic. The iron arsenide, $FeAs_2$, is not magnetic. No signs of Mn_2As have been observed. Arsenic is insoluble in sodium hydroxide.

Allotropic modifications.—When grey arsenic is quickly heated in a current of hydrogen, black glittering hexagonal crystals of arsenic are deposited nearest the hot portion of the tube; and, further on, a yellow powder is deposited. These are supposed to represent two allotropic modifications of arsenic. The specific gravity of the black crystalline variety, sometimes called *metallic arsenic*, is 5.7, and of the yellow 2.03. The yellow variety, called *α*-arsenic or yellow arsenic, is formed by the rapid condensation of arsenic vapour, and it is made by distilling arsenic in a current of carbon dioxide, and passing the vapour through a U-tube containing carbon disulphide in which the arsenic is condensed as it emerges from the hot tube by coming in contact with another stream of cold carbon dioxide; or by passing the vapour directly into a U-tube immersed in liquid air. Yellow arsenic is soluble in carbon disulphide from which it is deposited on evaporation in the form of rhombohedral crystals belonging to the cubic system. Yellow arsenic is very sensitive to light and quickly passes into the grey variety when exposed to light, and instantaneously under the influence of the magnesium light. Some yellow arsenic is formed when arsenious oxide is reduced by zinc dust in the presence of carbon disulphide. H. Erdmann assumed that yellow arsenic and arsenious oxide have a similar constitution:



but this is an unproved hypothesis. The effect of yellow arsenic on the boiling and freezing point of carbon disulphide corresponds with the formula As_4 . A reddish-brown variety of arsenic, specific gravity 3.764, is said to be deposited from carbon disulphide on long standing.¹ The so-called *brown arsenic* is formed by the action of oxygen or arsenious bromide on a carbon disulphide solution of yellow arsenic, or by the reduction of arsenic compounds; it is probably a variety of grey arsenic. The black modification of arsenic, called β -arsenic or black arsenic, is formed by the slow condensation of arsenic vapour. It is either a black powder, specific gravity 4.6 to 4.7, or a brittle glassy mass with a conchoidal fracture. At 360° it passes into the grey variety. Hence, by the rapid cooling of arsenic vapour, yellow arsenic is formed, and this passes by irreversible changes into black and then into the more stable grey variety; Yellow $As \rightarrow$ black $As \rightarrow$ grey As . It is possible that grey, brown, and metallic arsenic are all the same variety in different states of subdivision.

Atomic and molecular weights.—Numerous analyses of arsenic compounds show that the atomic weight ($O = 16$) lies somewhere between 74.90 and 75.22. The best representative value is taken to be 74.96; and the molecular weight of no known volatile compound of arsenic has less than 75 parts of arsenic per molecule when the molecular weight is determined by the vapour density (Avogadro's hypothesis). The specific heat of arsenic is 0.083. Hence by Dulong and Petit's rule, the atomic weight is nearly $6.4 \div 0.083 = 77$. Arsenic in the arsenates is isomorphous with phosphorus in the phosphates, and the same value is obtained for the atomic weight by Mitscherlich's rule. The vapour density of arsenic at about 860° , air = 1, is 10.2; and 5.45 at 1714° ; and 5.37 at 1736° . The molecular weight at the high temperatures thus conforms with a two-atom molecule and a molecular weight $5.4 \times 28.9 = 156$; and at low temperatures with a four-atom molecule.

History.—Arsenic was known to the ancients and considered by them to be a kind of sulphur. Aristotle mentions a substance, *σανδαράχη* (sandarache), which appears to have been red arsenic sulphide or realgar, and was called by Theophrastus *ἀρσενικόν* (arsenikon), meaning "potent"; the yellow sulphide was called *auripigmentum* and *arsenicum*. The former term was afterwards contracted to *orpiment*, a term which remains with us to-day. The element arsenic was prepared by Albertus Magnus, about 1250, and it was considered by the later alchemists to be a bastard or semi-metal. Brandt first showed that "white arsenic" is the calx of arsenic; and since the establishing of Lavoisier's theory of oxidation, white arsenic has been considered to be the oxide of the element.

Uses.—Arsenic is used in the manufacture of arsenic compounds—arsenic trioxide, etc., and in certain alloys. The presence of a trace in lead—1:1000—makes lead harder. "Chilled shot" is hardened with arsenic. The addition of arsenic lowers the melting point of the lead and increases its surface tension, so that when the shot is made by allowing the molten lead to drop from a height into water, the shot becomes spherical before it is cooled by the water.

¹ Arsenic suboxide, As_2O_3 , is said to be formed together with some of the allotropic modifications of arsenic when the element is sublimed in open tubes. There is some doubt about this.

§ 2. Antimony—Occurrence, Preparation, and Properties.

Atomic weight, $\text{Sb} = 120.2$; molecular weight, $\text{Sb}_4 = 480.8$. Ter-, quadri-, and quinquevalent. Melting point, 630.6° ; boiling point over 1300° .

Occurrence.—Antimony occurs free in small quantities in Borneo and a few other places. It is nearly always accompanied by some arsenic. Antimony occurs combined with oxygen as *antimony bloom*, Sb_2O_3 ; and as *antimony ochre*, Sb_2O_4 ; combined with sulphur as *stibnite* or *grey antimony ore*, Sb_2S_3 ; and as *antimony blende* or *red antimony*, $\text{Sb}_2\text{S}_2\text{O}$. It also occurs combined with sulphur and the metals.

Preparation.—Antimony is usually extracted from the native sulphides by heating the pulverized ore with scrap iron in a plumbago crucible. The iron combines with the sulphur forming a slag of iron sulphide which floats on the surface of the molten antimony: $\text{Sb}_2\text{S}_3 + 3\text{Fe} = 2\text{Sb} + 3\text{FeS}$. In another process, the crude sulphide is melted in such a way that the molten sulphide flows away from the less fusible rocky impurities. This process is called *liqumtion*.—The liquated sulphide is then mixed with about half its weight of charcoal and carefully roasted so as to convert the sulphide into oxide: $2\text{Sb}_2\text{S}_3 + 9\text{O}_2 = 2\text{Sb}_2\text{O}_3 + 6\text{SO}_2$. Part of the antimony oxide condenses in the flues, and a residue of Sb_2O_4 and unchanged sulphide remains behind. This is mixed with charcoal and sodium carbonate, and heated in a crucible. The reactions are taken to be: $\text{Sb}_2\text{O}_4 + 4\text{C} = 4\text{CO} + 2\text{Sb}$; and $3\text{Na}_2\text{CO}_3 + 6\text{C} + \text{Sb}_2\text{S}_3 = 9\text{CO} + 3\text{Na}_2\text{S} + 2\text{Sb}$. The antimony obtained by this process is subsequently refined by fusing it with a little nitre so as to oxidize the contaminating arsenic, lead, sulphur, etc. Antimony purified by two or three remeltings is the antimony of commerce.

Properties.—Antimony is a silvery-white solid with a high metallic lustre and a crystalline (rhombohedral) structure. It is very brittle and can be easily pulverized. Like the non-metals it is a poor conductor of heat, but it has a high specific gravity—6.7 to 6.8.

From its physical properties, antimony, like arsenic, would be classed with the metals, but its metallic characters are more pronounced than those of arsenic. Antimony melts at 629.2° in an atmosphere of carbon monoxide; and boils at 1440° . When the molten element is allowed to cool slowly and partially solidify in a crucible, the uncongealed portion may be poured off. The interior of

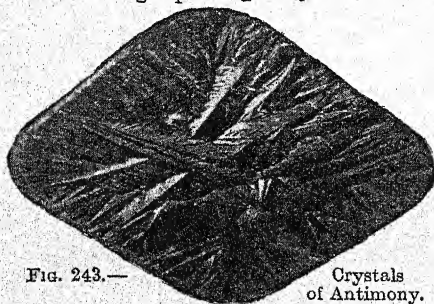


FIG. 243.—

Crystals
of Antimony.

the crucible is then lined with well-formed rhombohedral crystals of antimony isomorphous with arsenic. The reduced photograph, Fig. 243, illustrates the crystalline surface of a sample of 99.5 per cent. antimony from the St. Helen's Smelting Co. In the act of solidification lead contracts, but antimony expands slightly. Hence molten mixtures of antimony with other metals, when poured into moulds, take the fine and sharp impressions of the mould. The more important alloys of antimony are: *type metal*: lead, 75; tin, 5; antimony, 20. *Stereotype metal*: lead, 112; tin, 3; antimony, 18. *Britannia metal*: copper, 3; tin, 140; antimony, 7.

Antimony does not tarnish readily on exposure to dry air, but it is oxidized slowly by moist air. Antimony is used to cover other metals like brass and lead alloys. *Antimony black* is finely powdered antimony which is used to coat plaster casts, to make them imitate metals. When heated in air or oxygen, antimony burns with a bright bluish flame forming antimony trioxide, Sb_2O_3 . Antimony combines directly with the halogens. The action is vigorous, and the combining element becomes incandescent. With chlorine, antimony trichloride, SbCl_3 is formed. Antimony also unites with sulphur, phosphorus, and arsenic, forming sulphides, phosphides, and arsenides respectively. Dilute hydrochloric and sulphuric acids have little or no action upon antimony, but the more concentrated acids respectively form chloride: $2\text{Sb} + 6\text{HCl} = 2\text{SbCl}_3 + 3\text{H}_2$; and an unstable antimony sulphate: $2\text{Sb} + 6\text{H}_2\text{SO}_4 = 6\text{H}_2\text{O} + 3\text{SO}_2 + \text{Sb}_2(\text{SO}_4)_3$. If air be excluded, hydrochloric acid does not attack the element. Antimony thus behaves towards these acids like a typical metal. Dilute nitric acid has scarcely any action, but it probably forms an unstable antimony nitrate, $\text{Sb}(\text{NO}_3)_3$. Concentrated nitric acid does not dissolve the metal but rather oxidizes it to insoluble Sb_2O_3 or Sb_2O_5 , or a mixture of Sb_2O_4 and Sb_2O_5 .

Allotropic modifications.—Like phosphorus and arsenic, antimony exhibits allotropism. The variety now under discussion is called common or rhombohedral or β -antimony. Yellow or α -antimony is a metastable variety which is formed when antimony hydride gas, SbH_3 , is treated with air at -90° : $4\text{SbH}_3 + 3\text{O}_2 = 4\text{Sb} + 6\text{H}_2\text{O}$. This passes into black antimony on exposure to light. It is not clear whether or not black antimony is an intermediate form between α - and β -antimony. The metastable variety is said to be made by the rapid cooling of antimony vapour. Under these conditions, an amorphous black powder is obtained with a specific gravity 5.3. This variety slowly passes into rhombohedral antimony at 100° and rapidly at 400° . Gore (1855) found that if a current of electricity be passed through a solution of antimony trichloride in hydrochloric acid—using an antimony anode, and a platinum cathode—an amorphous powder of specific gravity 5.78 is deposited on the cathode. The cathode has then the appearance of a smooth polished graphite rod. The deposit appears to be solid solution of antimony trichloride in metastable or α -antimony. If this deposit be rubbed or scratched, an explosion occurs. The explosion is attended by the allotropic transformation of the metastable or α -form of antimony into the stable β -form or the rhombohedral variety, at the same time the temperature rises to about 250° , and 19.6 Cals. of heat are evolved per gram of antimony. Clouds of antimony trichloride are given off at the same time. Hence the term *explosive antimony* is applied to the solid solution of the trihalide in α -antimony. The evidence that explosive antimony is a solid solution turns on the fact that the halide is not removed by the solvent action of a mixture of alcohol and ether; and, when compressed at a low temperature and warmed gradually, the halide oozes out and condenses on the cold parts of the tube. The β -antimony, and that formed by explosion gives the same heat of reaction with bromine, and hence the two varieties are supposed to be the same.

Atomic and molecular weights.—The many analyses of antimony compounds show that the atomic weight of this element (oxygen = 16) lies

somewhere between 119.79 and 122.53; the best representative value is considered to be 120.2. The molecular weight of all known volatile compounds of antimony show that this number is the smallest weight of that element which enters into the composition of any one of its molecules. The vapour density of antimony at 2000° corresponds with the one-atom molecule; at 1640° very nearly with a two-atom molecule, and at 1440° , very nearly with the molecule Sb_3 . By analogy with arsenic and phosphorus, the latter number probably means that the vapour contains a mixture of Sb_4 and Sb_2 molecules. According to Dulong and Petit's rule, since the specific heat of antimony varies from 0.046 at 186° to 0.0537 at 300° , the approximate atomic weight of antimony varies from $6.4 \div 0.046 = 139$; to $6.4 \div 0.0537 = 119$. This agrees with the atomic weight 120.2.

History.—According to Dioscorides and Pliny, *stimon* and *stibium*—that is, stibnite or antimony sulphide—has long been employed by the women of the East as a medicine, and as an article of toilet for darkening the eyebrows. It is mentioned in this connection in the Old Testament (2 Kings ix. 30; Ezekiel xxiii. 40). Pliny terms it *stibium* and in a Latin translation of Geber, it is called *antimonium*. Both terms were in common use¹ up to the time of Lavoisier for antimony sulphide. The latter term is supposed to be derived from the Greek *ἀντί* (anti), against; *μόνος* (monos), one; or French *moine*, a monk. At the end of the eighteenth century, the properties of antimony had probably been investigated more carefully than the properties of any other element. The preparation of the element, and the known and imagined properties of antimony were described by Basil Valentine—a Benedictine monk of South Germany—about the fourteenth century. In his book entitled *Triumph-Wagen des Antimonii* (The Triumphant Chariot of Antimony), Valentine approaches the subject with awed devotion: "He who would write of antimony needs a great consideration and a most ample mind. . . . In a word, one man's life is too short to be perfectly acquainted with all its mysteries." The same remark might be applied with equal force to any one of the elements.

§ 3. Bismuth—Occurrence, Preparation, and Properties.

Atomic weight, Bi = 208; molecular weight, $\text{Bi}_2 = 416$. Ter-, quadri-, and quinquivalent. Melting point, 269° ; boiling point, 1435° .

Occurrence.—Bismuth is found in many localities—Bolivia, Saxony, etc.—in a fairly pure condition in a free state. It also occurs combined with sulphur as *bismuth glance*, Bi_2S_3 ; with tellurium as *tetradymite*, Bi_2Te_3 ; and with oxygen as *bismite* or *bismuth ochre*, Bi_2O_3 . *Bismuthite* is a hydrated carbonate. Most of the bismuth in commerce comes from Bolivia, and some from Australia and Bohemia. In the United States (Indiana) some bismuth is obtained as a by-product in the refining of lead ores.

Preparation.—The ore is first concentrated by washing away substances (impurities) of a lower specific gravity, or by passing the crushed ore

¹ Native antimony sulphide was called "kohl" by the Arabians, and later the name was changed to "alkol," and later still into "alcohol." In the Middle Ages, this latter term was applied to almost any fine powder produced by titration or sublimation, and later, distillation. The term "alcohol," at later periods, was gradually confined to "spirits of wine" produced by distillation.

through an intense magnetic field—magnetic concentration—to grade the mixture according to the magnetic susceptibility of the constituents. Bismuth is isolated by roasting the sulphide so as to form the oxide, $\text{Bi}_2\text{O}_3 : 2\text{Bi}_2\text{S}_3 + 9\text{O}_2 = 2\text{Bi}_2\text{O}_3 + 6\text{SO}_2$. The oxide is reduced to the metal by heating it with charcoal in crucibles, or in a reverberating furnace. If the metal is to be refined, it is fused on an inclined plate in contact with air. The molten bismuth is relatively slowly oxidized near its melting point, whereas the impurities are usually readily oxidized, or volatilized; and the molten bismuth can be run off from the oxides. The commercially "pure" bismuth averages 99.3 to 99.9 per cent. of bismuth. The remaining impurities are copper, lead, sulphur, silver, etc.

Properties.—Bismuth is a greyish-white solid resembling antimony, but it has also a faint reddish tinge. Bismuth is hard, brittle, lustrous, crystalline, and like antimony, expands about 2.3 per cent. in passing from the liquid to the solid condition. It is a very bad conductor of heat, possessing less than one-fiftieth the conductivity of silver. The metallic qualities of bismuth are far more pronounced than is the case with antimony and arsenic. In arsenic, the non-metallic properties predominate; in bismuth, the metallic qualities preponderate. The specific gravity of bismuth is near 9.8. Antimony has many properties intermediate between those of bismuth and arsenic. There are two enantiomorphic modifications $\alpha\text{-Bi} \rightleftharpoons \beta\text{-Bi}$ with a transition temperature at 75° . Bismuth melts at 269° . If the molten metal be allowed to partly solidify in a crucible, and the uncongealed fluid be poured off, the crucible will be found lined with rhombohedral crystals of the element. Bismuth boils at 1435° , and it distills in an atmosphere of hydrogen if heated over 1100° . Bismuth oxidizes superficially on exposure to air, but if heated in air, it burns, forming Bi_2O_3 . Bismuth decomposes steam at a red heat, but it is not affected by cold air-free water. It unites directly with the halogens. The halides of arsenic, antimony, and bismuth are isomorphous. It does not form a hydrogen compound, whereas antimony and arsenic form trihydrides. Bismuth is but slightly attacked by hydrochloric acid—hot or cold, dilute or concentrated; it is very sparingly soluble in hot sulphuric acid forming bismuth sulphate, $\text{Bi}_2(\text{SO}_4)_3$, and sulphur dioxide. Bismuth is readily attacked by dilute and concentrated nitric acid forming bismuth nitrate, $\text{Bi}(\text{NO}_3)_3$. The reaction is usually symbolized: $2\text{Bi} + 8\text{HNO}_3 = 2\text{Bi}(\text{NO}_3)_3 + 4\text{H}_2\text{O} + 2\text{NO}$. The nitrate dissolves in aqua regia forming bismuth chloride, BiCl_3 . Thus, nitric acid with bismuth gives a soluble nitrate, with antimony an insoluble oxide, and with arsenic, a soluble oxyacid.

Atomic and molecular weights.—The atomic weight of bismuth, oxygen = 16, lies between 207.9 and 210.8; the best representative value is taken as 208; and the atomic weight deducted from Dulong and Petit's rule, when the specific heat of bismuth at 18° is 0.0303 , and $6.4 \div 0.0303 = 211$. Hence, 208 is taken to represent the atomic weight of bismuth. The vapour density of bismuth between 1600° and 1700° corresponds with a mixture of molecules of Bi and Bi_2 ; and at 2000° the molecule is monatomic.

Uses.—Bismuth is used in the manufacture of alloys. The bismuth alloys are usually somewhat hard, and fusible. Some anti-friction metals contain about 0.25 per cent. of bismuth. Many of the "fusible alloys"

melt in warm water. For instance, *Newton's metal* (tin, 3; lead, 5; bismuth, 8) melts at 94.5° ; *Rose's fusible metal* (tin, 1; lead, 1; bismuth, 2) melts at 93.75° ; and *Wood's fusible metal* (tin, 1; lead, 2; cadmium, 1; bismuth, 4) melts at 60.5° —over 150° below the melting point of the most fusible metal. Fusible alloys, which melt at a low temperature, are used for making safety plugs in steam boilers; fuses in electrical connections; in fire alarms; and in automatic sprinklers for buildings so that if a fire breaks out, the heat fuses a plug of the water pipe and thus allows a rush of water from the main. The gas pipe which enters a building can also be fitted with a piece of fusible alloy so that if a fire breaks out, the alloy will melt, choke the gas pipe, and stop the flow of gas. Fireproof doors can also be kept open by fusible plugs which allow the doors to close automatically in the event of fire. The oxide or nitrate is used in making some kinds of optical glass, and in the decoration of pottery with lustres. The basic nitrate was once used as a cosmetic, but is now largely displaced by the cheaper zinc oxide. Some of the bismuth compounds are used medicinally.

History.—Metallic bismuth, called *marcasite*,¹ was described by Basil Valentine in the fourteenth century, and it was latter classed by Paracelsus as a bastard metal. Agricola considered it to be a true metal. Most of the writers about the 17th century confused bismuth with antimony (A. Libavius) and with zinc (N. Lemery). Paracelsus grouped together *marcasite*, zinc, and bismuth as bastard metals. According to his fantastic ideas, a "bastard metal" seems to have been one which played the rôle of impurity, contaminating the "primary matter of all metals." If these mixtures be "purged," said Paracelsus, "the bastard metal will in part separate or remain as *fæces*, and the formation of the pure metal can then take place." J. H. Pott (1739) first demonstrated the characteristic properties of bismuth; and its reactions were later studied by S. F. Geoffroy (1753), and by T. Bergmann (1780). The name bismuth is supposed to be derived from the German "*Weissmuth*," white matter.

§ 4. Arsenic and Antimony Hydrides.

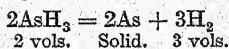
Preparation.—There are several reports of its formation of a *bismuth hydride*, but its existence has not been unequivocally demonstrated. A. Michaelis and A. Polis (1887) reported some bismuthines—*e.g.* $\text{Bi}(\text{CH}_3)_3$, $\text{Bi}(\text{C}_2\text{H}_5)_3$, and $\text{Bi}(\text{C}_6\text{H}_5)_3$ —Addition compounds of the first two compounds with bromine or with ethyl iodide, correspond with the ammonium derivatives. Triphenyl bismuthine forms addition compounds with chlorine and bromine—*e.g.* $\text{Bi}(\text{C}_6\text{H}_5)_3\text{Cl}_2$, and $\text{Bi}(\text{C}_6\text{H}_5)_3\text{Br}_2$ —which render the quinquevalency of bismuth as certain as that of antimony. The hydrides of arsenic—arsine, AsH_3 —and of antimony—stibine, SbH_3 —resemble phosphine and ammonia in composition and many properties. Arsenic is sometimes called *arseniuretted hydrogen*; and stibine, *antimoniuretted hydrogen*. Neither arsenic nor antimony unite directly with

¹ The term *marcasite* was formerly employed somewhat vaguely for any ore with a metallic appearance, and especially to ores now classed as pyrites. The term is now confined to a special variety of iron pyrites (*q.v.*).

hydrogen. The yellowish-brown powder obtained when arsenic is sublimed in hydrogen is not a hydride, but is arsenic itself. Hydrogen does not unite with arsenic under these conditions. Both hydrides are produced by the action of nascent hydrogen from metallic zinc or magnesium and sulphuric acid, upon solutions containing compounds of these two elements. Arsine is, and stibine is not, produced by the action of hydrogen generated when potassium hydroxide acts upon zinc. Fleitmann's test (1850) for distinguishing between arsenic and antimony is based upon this fact. When arsine or stibine is prepared by the action of nascent hydrogen upon an arsenic or an antimony salt, in an apparatus similar to that used for the preparation of hydrogen, the gas is accompanied by much hydrogen. Arsine practically free from hydrogen is made by the action of dilute sulphuric acid upon sodium arsenide or on zinc arsenide, Zn_3As_2 : $\text{Zn}_3\text{As}_2 + 3\text{H}_2\text{SO}_4 = 2\text{AsH}_3 + 3\text{ZnSO}_4$. Stibine is best made by the action of dilute sulphuric acid upon an alloy of antimony and zinc, or calcium, or magnesium.¹ The gas is passed through a U-tube immersed in liquid air. The stibine condenses to a white solid, and the hydrogen passes on. When the U-tube is removed from the liquid air, the stibine vaporizes.

Properties.—Both gases are very poisonous. One bubble of arsine is said to have produced fatal effects. A. F. Gehlen lost his life with this gas in 1815. H. S. Schulze more recently died from having worked with the gas without taking proper precautions. Stibine is less stable than arsine. Both gases when passed through a hot tube deposit the elements in the form of a metallic film. Both gases burn with a blue flame which give metallic films in contact with cold porcelain. Arsine gives arsenious oxide, As_2O_3 , on combustion; stibine, Sb_2O_3 . Stibine explodes when electric sparks are passed through the gas; arsine decomposes with the deposition of arsenic. Arsine can be exploded with fulminating mercury. Both compounds are endothermal: $\text{As} + 3\text{H} = \text{AsH}_3 - 36.7$ Cals.; and $\text{Sb} + 3\text{H} = \text{SbH}_3 - 81.8$ Cals. When an endothermal compound decomposes, heat is evolved. If such a compound begins to decompose at any point of its mass, the surrounding molecules are heated, and they too are decomposed developing more heat. The decomposition may thus traverse the whole mass if a vigorous enough impulse be imparted for the decomposition of a sufficient number of molecules to raise the temperature of the surrounding molecules to the temperature of decomposition. Concentrated sulphuric acid decomposes the gases, and therefore this liquid is not used for drying the gases. Calcium chloride, however, can be used. Arsine liquefies at -55° , and solidifies at -119° ; stibine liquefies at -18° , and solidifies at -88° .

Composition.—The composition of arsine or stibine can be determined by passing electric sparks through the gas, arsenic or antimony is deposited, and three volumes of hydrogen are formed. *E.g.*:



If, say, arsine be passed over hot weighed copper oxide and the resulting water and copper arsenide are weighed, every one part by weight of

¹ Zinc arsenide is made by heating metallic zinc and arsenic in a closed crucible so as to melt the mass. The calcium antimonide is made by fusing antimony, calcium chloride, and sodium in an iron crucible. The other antimonides are prepared in a similar manner.

hydrogen corresponds with 24.987 grams of arsenic. The atomic weight of arsenic by a previous experiment is 74.96. Hence, the formula for arsine is $(\text{AsH}_3)_n$, where n is to be determined; similarly, the formula of stibine is $(\text{SbH}_3)_n$. But the vapour density of arsine (hydrogen = 2) is 77.9; and of stibine, 123.6. These numbers correspond with the formulæ AsH_3 and SbH_3 .

Action on silver nitrate.—The two gases, arsine and stibine, are distinguished by their behaviour towards solutions of silver nitrate. With *concentrated* silver nitrate solutions, arsine gives a yellow compound of silver arsenide, AsAg_3 , with silver nitrate: *viz.*, $\text{Ag}_3\text{As} \cdot 3\text{AgNO}_3$, thus, $\text{AsH}_3 + 6\text{AgNO}_3 = \text{AsAg}_3 \cdot 3\text{AgNO}_3 + 3\text{HNO}_3$. This compound is decomposed by water, forming metallic silver and arsenious acid: $\text{AsAg}_3 \cdot 3\text{AgNO}_3 + 3\text{H}_2\text{O} = 6\text{Ag} + 3\text{HNO}_3 + \text{H}_3\text{AsO}_3$. This is the principle of H. Gutzzeit's test (1879). If a *dilute* solution of silver nitrate be used, silver is at once precipitated, A. W. Hofmann's test (1860): $\text{AsH}_3 + 6\text{AgNO}_3 + 3\text{H}_2\text{O} = 6\text{Ag} + 6\text{HNO}_3 + \text{H}_3\text{AsO}_3$. Stibine under similar conditions forms a precipitate of silver antimonide, Ag_3Sb ,—according to Lassargne's reaction: $\text{SbH}_3 + 3\text{AgNO}_3 = \text{Ag}_3\text{Sb} + 3\text{HNO}_3$. The precipitate is contaminated with some metallic silver formed during the action of the hydrogen on the nitrate. This reaction serves to distinguish and separate arsenic from antimony. The difference in the two reactions depends upon the fact that arsine is oxidized to arsenious acid more readily than the stibine, and thus arsine, like phosphine, acts as a reducing agent.

Marsh's test.—The underlying principle of the following test for arsenic was devised by J. Marsh in 1836. Place about 3 grams of metallic

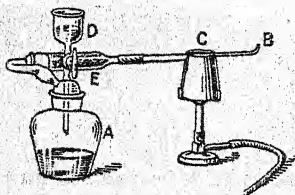


FIG. 244.—Modified Marsh's Apparatus.

zinc in a small flask A, Fig. 244, and add 20 c.c. of sulphuric acid (1 volume of acid, 8 volumes of water). The exit tube is fitted with a calcium chloride drying tube, E. When all the air has been expelled from the apparatus by the hydrogen, light the jet of gas issuing from the exit tube B. Light the gas burner so that the exit tube is heated at C. If no mirror is formed in the tube near C, the reagents are free from arsenic and antimony. Recharge the apparatus, and add the solution to be tested *via* the tap funnel D. If arsenic or antimony be present, a mirror will be deposited in the tube in 15 or 20 minutes. The gas is then extinguished and the exit tube disconnected. The mirror can be tested for arsenic by heating the mirror with a small flame while the tube is held in an inclined position. A garlic-like odour can be detected if 0.01 milligram of arsenic is present. When the tube is cold, arsenic trioxide will be found at a little distance from the flame in octahedral crystals which can be seen with the naked eye. These three results—mirror, garlic-like odour, and octahedral crystals—prove the presence of arsenic. Cumulative evidence is furnished by the application of other tests. If antimony be present, the mirror will be found nearer the hot flame C than is the case with arsenic, indeed, some antimony may be deposited in the tube before the gas reaches the hottest portion of the tube because stibine is more readily decomposed than arsine. If the exit tube be not heated at C, and the gases be burned

with hydrogen at the end of the tube, a piece of glazed porcelain held in the flame will receive a steel grey or black deposit of arsenic or a velvety brown or black deposit of antimony. If the deposit be treated with sodium hypochlorite solution, arsenic dissolves, while antimony remains insoluble. This illustrates the more ready oxidizability of arsenic than antimony.

Antimony forms but one hydride with hydrogen, SbH_3 ; arsenic forms a second hydride—arsenic dihydride—supposed to be As_2H_2 . This is the velvety-brown powder deposited when arsine is partially oxidized; when a jet of burning arsine impinges on porcelain; when potassium arsenide, K_3As , obtained by the reaction $3\text{KOH} + \text{AsH}_3 = \text{K}_3\text{As} + 3\text{H}_2\text{O}$, furnishes the solid hydride when treated with water: $2\text{K}_3\text{As} + 6\text{H}_2\text{O} = \text{As}_2\text{H}_2 + 6\text{KOH} + 2\text{H}_2$; when an electric discharge acts on arsine; and when water is electrolyzed with an arsenic cathode. The corresponding compounds of phosphorus and nitrogen are unknown. Cacodyl, $\text{As}_2(\text{CH}_3)_4$, is the only compound of arsenic which bears any analogy with P_2H_4 , and with hydrazine, N_2H_4 .

§ 5. Halogen Compounds of Arsenic, Antimony, and Bismuth.

Arsenic trichloride, AsCl_3 , is formed by the direct union of the elements when arsenic is brought in contact with chlorine; and it is also formed by distilling a mixture of arsenic trioxide with hydrochloric acid. The distillate of arsenic trichloride is a heavy, colourless, viscid liquid, boiling at 132° . The liquid freezes to white needle-like crystals at -18° . The mode of formation of arsenic trichloride: $\text{As}_4\text{O}_6 + 12\text{HCl} \rightleftharpoons 4\text{AsCl}_3 + 6\text{H}_2\text{O}$, suggests that arsenic trichloride is a salt and arsenious oxide a base. The basic properties of arsenious oxide are feeble, as is shown by the hydrolysis of the trichloride in contact with water. With water, arsenic trichloride forms arsenic oxychloride, $\text{AsCl}(\text{OH})_2$ or AsOCl ; with boiling water, arsenious acid is formed: $\text{AsCl}_3 + 3\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{AsO}_3 + 3\text{HCl}$, and the arsenious acid decomposes with the separation of arsenious oxide, As_4O_6 . There is some doubt if arsenic pentachloride, AsCl_5 , is formed, when arsenic trichloride is cooled with an excess of chlorine. Some consider the alleged arsenic pentachloride is a solution of chlorine in arsenic trichloride. The doubts arise from the fact that the freezing point curve of mixtures of chlorine and arsenic trichloride show no signs of a maximum corresponding with the compound in question. The pentachloride, if formed, decomposes at -28° : $\text{AsCl}_5 \rightleftharpoons \text{AsCl}_3 + \text{Cl}_2$. Arsenic trifluoride, AsF_3 , is formed as a volatile fuming liquid boiling at 63° when arsenious oxide, As_4O_6 , is distilled with the materials—calcium fluoride and sulphuric acid—used to generate hydrofluoric acid. Arsenic trifluoride is decomposed by water with the formation of hydrofluoric acid. Arsenic pentafluoride, AsF_5 , is known. Arsenic tribromide is a liquid which gives colourless prisms when cooled; arsenic triiodide forms orange-red rhombohedral crystals. These compounds are formed by the direct union of the elements. Arsenic pentiodide, AsI_5 , and also the iodides AsI and As_2I_4 (or AsI_2) have been reported, the first named is very doubtful.

Antimony trichloride, SbCl_3 , is formed by the action of dry chlorine on an excess of antimony, and also by the distillation of an intimate mixture of antimony or stibnite with mercuric chloride. It forms colourless crystals melting at 73° , and boiling at 223° . The vapour density

225.4 corresponds with the formula SbCl_3 . "Butter of antimony" is a viscid mass formed by the action of hydrochloric acid on stibnite. When chlorine is passed through fused antimony trichloride, antimony pentachloride, SbCl_5 , is formed as a colourless or slightly yellow liquid which crystallizes at -6° , fumes in air, and boils with partial dissociation into antimony trichloride and chlorine at 140° . Antimony pentachloride thus resembles phosphorus pentachloride: $\text{SbCl}_5 = \text{SbCl}_3 + \text{Cl}_2$. Antimony trichloride is decomposed by water forming the antimony oxychloride as a white precipitate. The composition of the oxychloride depends upon the amount and temperature of the water. Thus, with one part of antimony trichloride and 1.7 parts of water at ordinary temperatures, a precipitate of SbOCl is formed; with 5 to 50 parts of water, a precipitate of $2\text{SbOCl} \cdot \text{Sb}_2\text{O}_3$ is formed; and by boiling with an excess of water, Sb_2O_3 is precipitated. The oxychloride was once used in medicine under the name "powder of Algaroth"—after V. Algarotus. Antimony also forms crystalline trifluoride, triiodide, and tribromide by the direct action of the proper elements on one another.

Bismuth trichloride, BiCl_3 , is formed by the direct action of chlorine on bismuth; by heating bismuth with mercuric chloride; and by the action of *aqua regia* on bismuth. Bismuth trichloride is a white crystalline mass, deliquescent in air. It melts at 227° and boils at 428° . Its vapour density 328 corresponds with the formula BiCl_3 . Unlike arsenic and antimony oxychlorides, bismuth oxychloride, BiOCl , is not decomposed by water. When bismuth trichloride is dissolved in a little water it forms a syrupy liquid, but with an excess of water, a white precipitate of bismuth oxychloride, BiOCl . The dried precipitate is sometimes called "pearl white," and used as a white pigment. Bismuth oxychloride is insoluble, and antimony oxychloride is soluble in tartaric acid. The sparing solubility of bismuth oxychloride, its general appearance, and the fact that it turns grey on exposure to light recalls the properties of univalent silver and mercury halides. In order to emphasize this, it has been assumed that BiO acts as a univalent radicle called bismuthyl, so that BiOCl is bismuthyl chloride. Bismuth tribromide, BiBr_3 , is an orange yellow solid; bismuth triiodide, BiI_3 , a reddish-brown powder; and bismuth trifluoride, BiF_3 , a white solid. These compounds are prepared in a similar manner to the trichloride. They too give precipitates of the basic salts with water. Bismuth triiodide unites directly with hydrogen iodide forming a monobasic acid— $\text{HBiI}_4 \cdot 4\text{H}_2\text{O}$, called hydr-iodo-bismuthous acid. The potassium iodobismuthite, KBiI_4 , crystallizes in red plates.

§ 6. Arsenic Trioxide, Arsenious Acids.

The molecule of arsenic trioxide or arsenious oxide is often symbolized As_4O_6 , although usually written: As_2O_3 . This oxide is perhaps the most important compound of arsenic, and in commerce it is often called "white arsenic," or simply "arsenic."¹ Small quantities occur free in nature. Arsenic trioxide is formed by oxidizing arsenic with nitric acid and by the

¹ A little known arsenic suboxide, As_2O , is said to be formed together with allotropic modifications of the element when metallic arsenic is sublimed in an open tube.

combustion of arsenic. Unlike phosphorus, the oxidation of burning arsenic stops when the trioxide is formed, phosphorus passes a stage further and forms the pentoxide. Commercial "white arsenic" is a by-product in the roasting of arsenical ores, the arsenic is oxidized to the trioxide and the fumes are led through a series of chambers or flues arranged to present an extended condensing surface to the gases. The crude product—"arsenical soot"—is collected periodically. It is refined by sublimation from iron cylinders.

Properties.—Like phosphorous oxide, arsenious oxide occurs in several different forms. Amorphous arsenic trioxide, or vitreous arsenic, is a colourless, non-crystalline, glassy substance produced when the vapour of arsenic trioxide is *slowly* condensed at a temperature slightly below its vaporizing temperature. It melts at about 200° , and volatilizes at 218° . 100 parts of water at 10° dissolve 3.67 parts of vitreous arsenic oxide in 6 hours, the solubility diminishes on standing. *E.g.* after standing 1 day, the solubility fell to 3.31; 21 days, 1.71; and $2\frac{1}{2}$ years, 1.71. Vitreous arsenic trioxide can be preserved unchanged in a sealed tube, but if it be exposed to the air it gradually becomes opaque and forms crystals of octahedral arsenic trioxide belonging to the cubic system. Similar crystals are produced when the vapour of arsenic trioxide is *rapidly* condensed, and when aqueous or hydrochloric acid solutions of the trioxide are allowed to crystallize. Octahedral arsenic trioxide has less than one-third the solubility of the vitreous variety. If a hot saturated solution of arsenic trioxide in water be allowed to cool slowly in a dark room, a flash of light is produced as each crystal is formed, while the effect on shaking the crystallizing solution has been likened to a "fine display of phosphorescent light at sea." The specific gravity of vitreous arsenic trioxide is 3.74, and of the octahedral form, 3.63. The passage from the vitreous to the octahedral variety is attended by an evolution of heat, 5.3 Cals. Octahedral arsenic trioxide vaporizes without fusion, but if heated under pressure it melts and forms the vitreous variety.

If a hot saturated solution of arsenic trioxide in potassium hydroxide be cooled, prismatic needle-like crystals of rhombic arsenic trioxide separate. The latter are sometimes said to belong to the monoclinic system—see "Antimony Trioxide." This variety is also formed when either of the preceding forms of arsenic trioxide are heated for a long time at 200° . Its specific gravity is 4.15. If the lower part of a sealed tube containing arsenic trioxide be heated above 400° , the lower part will contain vitreous, the middle prismatic (rhombic), and the upper part octahedral (cubic) arsenic trioxide. This experiment is due to H. Debray (1864). A mixture of the prismatic and octahedral crystals will be found in the zone between the octahedral and prismatic crystals.

Fig. 245 will give some idea of the crystals in the intermediate zone as they appear magnified under a $1\frac{1}{2}$ " objective. The vitreous form is the unstable

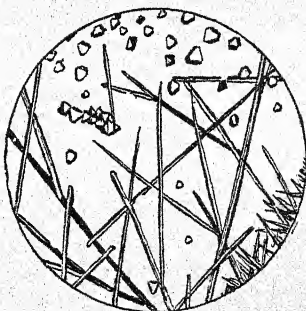


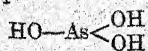
Fig. 245.—Dimorphism of Arsenic Trioxide.

variety, and the octahedral form the stable variety at ordinary temperatures. The vitreous variety can exist at ordinary temperatures because of the slow speed of transformation into the octahedral form. At 100° , the speed is quite fast. At higher temperatures still, the octahedral form passes into the monoclinic variety. Pressure alone (2500 kgms. per sq. cm.) at 150° will transform arsenic trioxide into the vitreous variety.

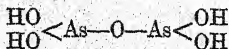
All three varieties of arsenic trioxide vaporize at 218° . The vapour density between 500° and 800° corresponds with the molecule As_2O_3 ; and at 1732° , with As_2O_3 . Arsenic trioxide is quite stable if heated in air or oxygen at 100° . Oxidizing agents—like iodine, hydrogen peroxide, nitric acid, and ozone—transform it into arsenic acid. Arsenic trioxide dissolves in hydrochloric acid forming arsenic trichloride. Reducing agents—stannous chloride—transform it into arsenic and arsine. Arsenic trioxide is very poisonous—0.06 gram is near the fatal dose for an ordinary man; but the habitual use of small doses makes the system more or less immune to the effects of much larger quantities.¹ The antidote is freshly precipitated ferric hydroxide made by adding aqueous ammonia to a ferric salt—e.g. ferric chloride. The efficacy of the antidote depends on the formation of an insoluble compound with iron hydroxide, or else on the fixation of the soluble arsenic by adsorption on the ferric hydroxide. Freshly prepared hydroxide is the most efficacious.

Uses.—Arsenic trioxide is used in the manufacture of pigments—*Scheele's green*, for instance, is a copper hydrogen arsenite, HCuAsO_3 , once largely used to colour wall-paper. Several cases of arsenical poisoning are believed to have been produced by the arsenical dust from such papers. *Paris green* or *Schweinfurt green* is a double salt of copper acetate and copper meta-arsenite— $3\text{Cu}(\text{AsO}_2)_2 \cdot \text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2$. It is used as an insecticide, and for poisoning vermin. Arsenic trioxide is also used in preserving the skins of animals, and in the manufacture of glass and opaque enamels. The sulphides, realgar and orpiment, are used as paints and pigments; the salts are used medicinally, in dyeing, and in calico printing.

Arsenites.—Aqueous solutions of arsenious oxide exhibit a feebly acid reaction, and redden blue litmus, probably due to the formation of an unstable arsenious acid, H_3AsO_3 . It differs from the corresponding dibasic phosphorous acid in being tribasic, as is well illustrated by the silver salts Ag_3AsO_3 and Ag_2HPO_3 . This acid has never been isolated because the solution, on concentration, deposits crystals of the trioxide. Salts corresponding to ortho-, pyro-, and meta-arsenious acid are known. The constitution of the salts of these three acids may be understood from the graphic formulæ:



Ortho-arsenious acid.



Pyro-arsenious acid.



Meta-arsenious acid.

The salts are formed by precipitation or crystallization from solutions; not by ignition as in the case of the corresponding phosphorus acids.

¹ With reference to the Styrian arsenic eaters, J. F. W. Johnston (1855) says: "Arsenic is consumed chiefly for two purposes—First, to give plumpness to the figure, cleanness and softness to the skin, and beauty and freshness to the complexion. Secondly, to improve the breathing and give longness of wind, so that steep and continuous heights may be climbed without difficulty and exhaustion of breath. Both these results are described as following almost exactly from the prolonged use of arsenic either by men or animals."

ACIDS.

Ortho-arsenious acid, $\text{As}(\text{OH})_3$,
 or H_3AsO_3
 Pyro-arsenious acid, $\text{As}_2\text{O}(\text{OH})_4$,
 or $\text{H}_4\text{As}_2\text{O}_5$
 Meta-arsenious acid, $\text{AsO}(\text{OH})$,
 or HAsO_2

SALTS.

Silver ortho-arsenite, Ag_3AsO_3
 Copper hydrogen arsenite, CuHAsO_3
 Calcium pyro-arsenite, $\text{Ca}_2\text{As}_2\text{O}_5$
 Ammonium pyro-arsenite, $(\text{NH}_4)_4\text{As}_2\text{O}_5$
 Potassium meta-arsenite, KAsO_2
 Lead meta-arsenite, $\text{Pb}(\text{AsO}_2)_2$

All the arsenites, except the alkali arsenites, are insoluble in water, and they are decomposed by carbonic and hydrosulphuric acids. The soluble arsenites react alkaline owing to hydrolysis. Arsenious acid and its salts are reducing agents, but not so powerful as phosphorous acid and the phosphites. Arsenious acid precipitates red cuprous oxide from an alkaline cupric solution. When heated, the arsenites are converted into arsenates and free arsenic; and when heated with charcoal, both the arsenates and the arsenites are reduced to elemental arsenic.

§ 7. Antimony Trioxide and Antimonious Acid.

Antimony trioxide or antimonious oxide is made by burning antimony in air, and by adding hot water to a solution of either antimony trichloride or antimony sulphate: $4\text{SbCl}_3 + 6\text{H}_2\text{O} = \text{Sb}_4\text{O}_6 + 12\text{HCl}$. The precipitated oxide is washed with a solution of sodium carbonate to remove the free acid, and finally with water.

Antimonious oxide is a white powder. It volatilizes just over 1500° , and its vapour density then corresponds with Sb_4O_6 . The vapour condenses in two distinct forms—prismatic needle-like crystals belonging to the rhombic system, and in regular octahedra belonging to the cubic system. These crystals are isodimorphous with the corresponding crystals of arsenious oxide. The octahedra are deposited nearest to the hot zone, and the prismatic crystals further along in the cooler part. A mixture of both will be found in between. The two forms occur in nature as the rhombic mineral *senarmonite*, and the cubic mineral *valentinite*. The latter crystallizes in well-formed octahedra. The former is the variety stable at high temperatures, the latter at low ones.

This oxide is very sparingly soluble in water, and the solution has no action on litmus. The oxide is insoluble in nitric and sulphuric acids, but it dissolves in hydrochloric acid forming antimony trichloride, SbCl_3 ; it also dissolves in alkalis. The basic qualities of antimony trioxide are also exemplified in its uniting with sulphuric and nitric acids to form salts: antimony sulphate, $\text{Sb}_2(\text{SO}_4)_3$; antimony nitrate, $\text{Sb}(\text{NO}_3)_3$. The oxide is also readily soluble in an aqueous solution of tartaric acid, $\text{H}_2(\text{C}_4\text{H}_4\text{O}_6)$, and in a boiling solution of acid potassium tartrate, $\text{HK}(\text{C}_4\text{H}_4\text{O}_6)$. The solution furnishes crystals of tartar emetic, or potassium antimonyl tartrate, $2[(\text{SbO})\text{K}(\text{C}_4\text{H}_4\text{O}_6)]\text{H}_2\text{O}$. This is the most important salt of antimony in commerce. If hydrochloric acid be added to an aqueous solution of tartar emetic, antimonyl chloride, SbOCl , that is antimony oxychloride, is formed; and if sulphuric or nitric acid be added, ortho-antimonious acid, $\text{Sb}(\text{OH})_3$, is precipitated because the antimony sulphate or nitrate first formed is immediately hydrolyzed: $\text{SbO}(\text{NO}_3) + 2\text{H}_2\text{O} \rightleftharpoons \text{HNO}_3 + \text{Sb}(\text{OH})_3$.

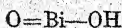
Antimonites and antimonious acids.—Ortho-antimonious acid is prepared by the action of acids upon tartar emetic as indicated above.

The corresponding arsenic compound has not been isolated, probably because arsenious hydroxide is not so basic as the corresponding antimony compound. Ortho-antimonious acid forms a white powder when dried at 100° . This hydrate readily loses one molecule of water, forming meta-antimonious acid, SbO.OH , or HSbO_2 . Meta-antimonious acid shows its acidic qualities by forming salts—meta-antimonites—when it is dissolved in alkalis; the hydrogen of the acid is replaced by the basic radicle, e.g. sodium meta-antimonite, $\text{NaSbO}_3 \cdot 3\text{H}_2\text{O}$, has been obtained in tetragonal crystals. It is almost insoluble in water. The graphic formulæ resemble those of the arsenious acids.

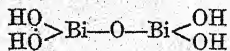
On boiling antimony trioxide with sodium hydroxide, a solution of sodium ortho-antimonite, Sb(ONa)_3 or Na_3SbO_3 , is formed. The salt is hydrolyzed when the solution is diluted with water: $\text{Sb(ONa)}_3 + \text{H}_2\text{O} = 2\text{NaOH} + \text{NaSbO}_2$, and also into Sb_2O_3 if much water be present. The filtered solution deposits crystals of sodium meta-antimonite on cooling. The latter is decomposed on further dilution into antimony trioxide. Hence, if sodium hydroxide or carbonate be added to antimony trichloride an almost quantitative precipitation of the trioxide, Sb_2O_3 , occurs. SbO in SbO.H , like KO in K.O.H , behaves as a univalent basic radicle and it is called antimonyl—e.g. antimonyl nitrate, $\text{SbO(NO}_3)_2$; and antimonyl sulphate $(\text{SbO})_2\text{SO}_4$. All these compounds are readily hydrolyzed by water and form acid and the trioxide, so that they cannot be made in aqueous solution. Pyro-antimonious acid has been reported— $\text{Sb}_2\text{O(OH)}_4$, or $\text{H}_4\text{Sb}_2\text{O}_7$, but there is some doubt if it has yet been isolated. The constitution of the antimonites will appear by comparison with the graphic formulæ for the corresponding arsenites.

8. Bismuth Trioxide.

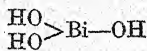
Bismuth trioxide is formed when bismuth is heated in air or when the hydrated oxide, carbonate, or nitrate is calcined. Bismuth trioxide is a cream-coloured powder which is not acted on by water. It is the most stable of all the three bismuth oxides. All the three known bismuth oxides form the trioxide when heated in air. Bismuth trioxide melts at 700° . It is dimorphous, and isodimorphous with the corresponding oxides of antimony and arsenic.¹ Bismuth trioxide is said to form three hydrates: $\text{Bi}_2\text{O}_3 \cdot \text{H}_2\text{O}$; $\text{Bi}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$; and $\text{Bi}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$. The possible graphic formula are respectively:



Metahydrate.



Mesohydrate.



Orthohydrate.

These hydrates have no acidic properties and do not combine with bases to form salts. The trihydrate is formed by pouring an acid solution of bismuth nitrate into an excess of aqueous ammonia: $2\text{Bi(NO}_3)_3 + 6\text{NH}_4\text{OH} = 6\text{NH}_4\text{NO}_3 + \text{Bi}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$. This at 100° passes into the monohydrate, $\text{Bi}_2\text{O}_3 \cdot \text{H}_2\text{O}$. Bismuth trioxide exhibits marked basic properties,

¹ A little known bismuth suboxide, BiO , or Bi_2O_2 , is said to be formed as a black precipitate when a bismuth salt is reduced by an alkaline solution of stannous chloride. It is possible that the precipitate is a mixture of metallic bismuth and of the hydrated oxide.

for it dissolves in acids to form salts—bismuth nitrate, $\text{Bi}(\text{NO}_3)_3$; bismuth sulphate, $\text{Bi}_2(\text{SO}_4)_3$, etc. With a small quantity of hydrochloric acid, it first forms bismuth oxychloride, BiOCl , and with an excess of acid, bismuth trichloride, BiCl_3 . Water converts all these salts into insoluble basic salts and free acid. The small amount of acid formed is practically without action on the precipitate at the extreme dilution, and precipitation is therefore practically complete: $\text{Bi}(\text{NO}_3)_3 + 2\text{H}_2\text{O} = 2\text{HNO}_3 + \text{BiO}(\text{NO}_3)$. H_2O . One part of bismuth nitrate in contact with 50,000 parts of water can be detected by the resulting turbidity. Bismuth oxynitrate, bismuthyl nitrate, $\text{BiO}(\text{NO}_3) \cdot \text{H}_2\text{O}$, or $\text{Bi}(\text{NO}_3)(\text{OH})_2$, is used medicinally and in the preparation of cosmetics. It was once called "magister of bismuth" or "Spanish white." The latter term is sometimes used commercially for a special grade of whiting. Since the bismuth oxynitrate becomes yellow on exposure to air, and blackens in the presence of hydrogen sulphide (e.g. the waters of Harrogate Springs), its use as a cosmetic has given way in favour of zinc oxide.

§ 9. Arsenic Pentoxide, and Arsenic Acids.

Arsenic pentoxide cannot be prepared, as in the case of phosphorus pentoxide, by the combustion of arsenic in oxygen gas because the product with arsenic is the trioxide. Arsenic acid, H_3AsO_4 , is made by digesting arsenic trioxide with nitric acid. The cold concentrated solution deposits rhombic crystals of $2\text{H}_3\text{AsO}_4 \cdot \text{H}_2\text{O}$ melting at about 36° , when heated to 100° , the crystals lose their water of crystallization and a crystalline powder once called ortho-arsenic acid, H_3AsO_4 , remains. It was also stated that if ortho-arsenic acid be heated to about 140° or 180° , pyro-arsenic acid, $\text{H}_4\text{As}_2\text{O}_7$, is formed: $2\text{H}_3\text{AsO}_4 = \text{H}_2\text{O} + \text{H}_4\text{As}_2\text{O}_7$; and when this is heated to about 200° it passes into meta-arsenic acid, HAsO_3 , by the loss of another molecule of water: $\text{H}_4\text{As}_2\text{O}_7 = 2\text{HAsO}_3 + \text{H}_2\text{O}$. As a matter of fact, the existence of only two compounds of arsenic pentoxide with water have been definitely established—viz., As_2O_5 , $4\text{H}_2\text{O}$, and $3\text{As}_2\text{O}_5 \cdot 5\text{H}_2\text{O}$. When arsenic acid is heated to 155° for about 14 days, to 170° for about 5 days, or to 210° for $2\frac{1}{2}$ hours, arsenic pentoxide is obtained as a white deliquescent glassy solid: $2\text{H}_3\text{AsO}_4 = 3\text{H}_2\text{O} + \text{As}_2\text{O}_5$. Unlike phosphorus pentoxide, if arsenic pentoxide be heated to a higher temperature—over 440° —it breaks down into the trioxide and oxygen: $2\text{As}_2\text{O}_5 = \text{As}_4\text{O}_6 + 2\text{O}_2$. Nitrogen pentoxide is also decomposed on heating. Like phosphorus pentoxide, arsenic pentoxide dissolves readily in water forming tribasic arsenic acid, H_3AsO_4 —ortho-arsenic acid. The pyro- and meta-acids are said to resemble the corresponding phosphoric acids, but they are less stable. Metaphosphoric acid is stable at high temperatures, but meta-arsenic acid is decomposed as indicated above. Both the pyro- and meta-arsenic acids form crystalline solids which dissolve in water with the evolution of heat, and the formation of the ortho-acid. Hence, unlike the corresponding phosphoric acids, aqueous solutions of meta- and pyro-arsenic acids cannot exist. Each of the arsenic acids forms salts isomorphous with the corresponding phosphate, but the salts of the meta- and pyro-acids, when dissolved in water, pass into the ortho-compounds. If the nomenclature indicated when discussing the periodic acids were consistently followed, the ortho-acid would be $\text{As}(\text{OH})_5$. The term ortho- is applied to

the $\text{AsO}(\text{OH})_3$ acid for the same reasons that $\text{PO}(\text{OH})_3$ is called orthophosphoric acid. The soluble arsenates have a slightly alkaline reaction. When arsenic acid is treated with sodium hydroxide, with methyl orange as indicator, it behaves like a monobasic acid; and with phenolphthalein as indicator, as a dibasic acid. Hence, one reason why the arbitrariness of litmus as a standard for neutrality was emphasized on p. 171. The arsenates give a reddish-brown precipitate of silver arsenate, AgAsO_4 , when neutral solutions are treated with silver nitrate. The precipitate is analogous with yellow silver phosphate, Ag_3PO_4 ; arsenites in neutral solutions give a yellow precipitate under similar conditions.

§ 10. Antimony Pentoxide and Antimonic Acids.

Orthoantimonic acid, H_3SbO_4 , or $\text{O}=\text{Sb}(\text{OH})_3$. This acid is made by oxidizing antimonious chloride with a concentrated solution of nitric acid, and then diluting the solution with water. A white powder, approximating $\text{Sb}_2\text{O}_5 \cdot 6\text{H}_2\text{O}$, is formed. The freshly prepared hydrated oxide, $\text{Sb}_2\text{O}_5 \cdot 6\text{H}_2\text{O}$, dissolves more quickly than the aged hydrate. The solution deposits a white powder, approximating $\text{Sb}_2\text{O}_5 \cdot 4\text{H}_2\text{O}$, on standing for some time. Both hydrates pass into the ortho-acid, H_3SbO_4 , after standing over concentrated sulphuric acid in a desiccator for some time. At 200° , the ortho-acid passes into the pyro-acid. The corresponding salts are not known.

Pyro-antimonic acid, $\text{H}_4\text{Sb}_2\text{O}_7$, can be prepared as just stated; and also by the action of nitric acid on antimony at 100° . The acid obtained by the latter reaction is not constant in composition, for it is probably a mixture of different antimonic acids. The pyro-acid can be most easily made by the action of hot water on antimony pentachloride: $2\text{SbCl}_5 + 7\text{H}_2\text{O} = 10\text{HCl} + \text{H}_4\text{Sb}_2\text{O}_7$. The precipitate is dried at 100° . The pyro-acid, at 306° , is said to lose water and pass into meta-antimonic acid, HSbO_3 —sometimes called antimonic acid—pyro-antimonic acid, by the way, is sometimes erroneously called "meta-antimonic acid." The graphic formulæ of the arsenic and antimonic acids probably resemble those of the corresponding phosphoric acids.

Naming the antimonic acids.—There is some confusion in the nomenclature of the antimonic acids. Fashion has been stronger than system. Strictly, the unknown acid, $\text{Sb}(\text{OH})_3$, should be called "ortho-antimonic acid," but, as in the case of the phosphoric acids, the first dehydration product stable at ordinary temperatures, H_3SbO_4 , is called "ortho-antimonic acid." In conformity with the notation adopted for the phosphoric and arsenic acids, the next dehydration product, $\text{H}_4\text{Sb}_2\text{O}_7$, is called "pyro-antimonic acid," and the next dehydration product, HSbO_3 , "meta-antimonic acid." E. Fremy, who first made sodium pyroantimonate, $\text{Na}_2\text{H}_2\text{Sb}_2\text{O}_7 \cdot 6\text{H}_2\text{O}$, called it "sodium meta-antimonate," and the corresponding acid, $\text{H}_4\text{Sb}_2\text{O}_7$, "meta-antimonic acid," to distinguish it from "antimonic acid." It is best to keep to the uniform system already in use for the arsenic and phosphoric acids.

Antimony pentoxide, Sb_2O_5 .—When antimonic acid is heated to between 300° and 400° , antimony pentoxide, Sb_2O_5 , is formed; this begins to decompose at 440° , giving antimony tetroxide, Sb_2O_4 , the decomposition is complete at 750° or 800° . Antimony pentoxide is a straw yellow powder, almost insoluble in water. When moistened with water, however, the moist mass reddens blue litmus. Antimony pentoxide dissolves in concentrated hydrochloric acid forming a liquid which has oxidizing properties.

For instance, it decomposes potassium iodide with the separation of iodine: $\text{Sb}_2\text{O}_5 + 4\text{KI} + 10\text{HCl} = 2\text{SbCl}_3 + 4\text{KCl} + 2\text{I}_2 + 5\text{H}_2\text{O}$. If the meta-acid be allowed to remain for a long time in contact with water, it forms ortho-antimonic acid, H_3SbO_4 .

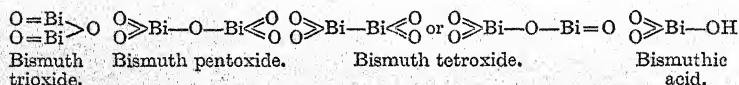
Antimoniates.—All the antimonic acids give antimonates when treated with an aqueous solution of potassium hydroxide. The normal ortho-antimonates have not been prepared, but primary salts are known. The oxide Sb_2O_5 can be regarded as antimonous orthoantimonate, $\text{Sb}(\text{SbO}_4)$. Numerous salts of pyro- and meta-antimonic acids have been prepared. Potassium pyro-antimonate, $\text{K}_4\text{Sb}_2\text{O}_7$, is best made by adding powdered antimony in small quantities at a time to four times its weight of fused nitre, KNO_3 . The fused mass presumably contains potassium ortho-antimonate, K_3SbO_5 , but when the cold mass is washed with water, the ortho-salt is decomposed forming potassium pyro-antimonate, $\text{K}_4\text{Sb}_2\text{O}_7$, and the addition of more water, or by rapid boiling, the salt is gradually changed into the acid salt: potassium dihydrogen pyro-antimonate, $\text{K}_2\text{H}_2\text{Sb}_2\text{O}_7 \cdot 6\text{H}_2\text{O}$, as granular powder sparingly soluble in cold water, but fairly soluble in hot water. The solution of this salt is used as a test for sodium salts with which it gives a precipitate of sodium dihydrogen pyro-antimonate, $\text{Na}_2\text{H}_2\text{Sb}_2\text{O}_7 \cdot 6\text{H}_2\text{O}$. On boiling potassium dihydrogen pyro-antimonate for a long time with an excess of water, it gradually passes into potassium dihydrogen ortho-antimonate, $\text{O} = \text{Sb}(\text{OK})(\text{OH})_2$, which is obtained as a gelatinous precipitate $2\text{KH}_2\text{SbO}_4 \cdot \text{H}_2\text{O}$ on evaporation of the solution. By further boiling, ortho-antimonic acid is formed. All the antimonates are decomposed by acids giving antimony pentoxide.

Antimony tetroxide, Sb_2O_4 .—When antimony trioxide is heated in air, it appears to smoulder, forming antimony tetroxide, Sb_2O_4 , which is yellow when hot, and white when cold. The powder is almost insoluble in water. The same oxide is formed when antimony pentoxide is heated to a red heat; the dissociation pressure is about 950° . Some consider the tetroxide to be antimonyl metantimonate, $\text{SbO} \cdot \text{SbO}_3$; others, antimony ortho-antimonate, $\text{Sb} \cdot \text{SbO}_4$. It forms salts, e.g. potassium hypoantimonate, $\text{K}_2\text{Sb}_2\text{O}_5$, or $\text{K}_2\text{O} \cdot \text{Sb}_2\text{O}_4$, when fused with potash. Hence, antimony tetroxide is generally considered to be the anhydride of an unknown acid, hypoantimonic acid— $\text{H}_2\text{Sb}_2\text{O}_5$. It is used as an oxidizing agent.

§ 11. Bismuth Pentoxide. Bismuthic Acids.

If chlorine be passed into a boiling aqueous solution of potassium hydroxide in which bismuth trioxide is suspended, a dark chocolate-brown precipitate is formed: $\text{Bi}_2\text{O}_3 + \text{KOC}l = \text{KCl} + \text{Bi}_2\text{O}_4$. The precipitate is washed with water, dilute nitric acid, and dried at 180° . The precipitate appears to be a mixture of bismuth pentoxide, Bi_2O_5 , with bismuth tetroxide, $\text{Bi}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$. The greater the excess and concentration of the lye the greater the yield of tetroxide. The tetroxide, more or less contaminated with other oxides, is formed when the trioxide is oxidized with potassium ferricyanide in alkaline solutions. In some respects the oxide, Bi_2O_4 , resembles antimony tetroxide; but the mode of formation and the general properties of the tetroxide resemble those of lead dioxide. When heated above 200° , bismuth tetroxide decomposes into the trioxide, Bi_2O_3 . If

the current of chlorine gas be continued until the precipitate becomes scarlet red, the washed precipitate has approximately the composition KBiO_3 —potassium metabismuthate, corresponding with metabismuthic acid, HBiO_3 . This appears to be rather a feeble acid. The salts are decomposed by water. Potassium bismuthate, KBiO_3 , is also formed as a dark red deposit on the anode when an almost boiling hot solution of potassium hydroxide (specific gravity 1.43) and potassium chloride in which bismuth trioxide is suspended, is electrolysed in a platinum dish. If the precipitate be washed and boiled for a short time in dilute nitric acid, scarlet red meta-bismuthic acid is obtained. The ortho-bismuthates are not known, but the oxide Bi_2O_4 can be regarded as bismuthous orthobismuthate, $\text{Bi}(\text{BiO}_4)$. When heated between 100° and 120° , bismuthic acid loses its water and some oxygen, and is said to form a little bismuth pentoxide, Bi_2O_5 , and the colour changes from red to dark brown. When heated to a higher temperature, bismuth trioxide is formed. The dark brown mass suddenly turns deep red at about 300° owing to the formation of a third modification of ordinary bismuth trioxide. Bismuth tetroxide is sometimes considered to be a basic salt of bismuthic acid, $\text{BiO} \cdot \text{BiO}_3$, where BiO is supposed to be a monad radicle. The graphic formula usually assigned to the bismuth oxides are :



Bismuth pentoxide is an unstable brown powder formed as indicated above. With hydrochloric, and the oxy-acids, bismuth pentoxide furnishes bismuth salts and chlorine or oxygen, thus behaving like antimony pentoxide, and both behave as if they were basic peroxides. If a solution of the so-called bismuth pentoxide in 60 per cent. hydrochloric acid be added drop by drop to ice-cold *N*-sodium hydroxide, the yellow precipitate contained 95 to 98 per cent. of its bismuth in the quinquevalent form, and the equivalent of one molecule of sodium hydroxide. When neutralized with ice-cold 15 per cent. nitric acid and washed rapidly by decantation, a reddish-brown precipitate containing the equivalent of 98 per cent. of bismuth pentoxide is obtained, which when treated with 15 per cent. sodium hydroxide furnishes yellow sodium bismuthate. This shows that the higher oxides of bismuth have acidic properties, although the red product formed by the oxidation of bismuth trioxide in the presence of potassium hydroxide as discussed above has no acidic properties.

§ 12. The Sulphides of Arsenic, Antimony, and Bismuth.

Arsenic trisulphide, As_2S_3 .—This sulphide occurs in nature as the mineral *orpiment*—a term derived from the Latin *auri pigmentum*, golden paint. It is employed as a pigment under the name *King's yellow*. Arsenic trisulphide is formed when powdered arsenic and sulphur are heated together in the proportions indicated by the formula, As_2S_3 ; and it is also precipitated as a canary-yellow solid when hydrogen sulphide is passed into a solution of an arsenious salt acidified with hydrochloric acid (p. 757). The same remarks apply, *mutatis mutandis*, to antimony trisulphide,

Sb_2S_3 , precipitated as a brick-red powder; and to bismuth trisulphide, Bi_2S_3 , precipitated as a dark brown powder. These sulphides can be fused, and on cooling a brittle crystalline mass is obtained. When heated in air, they are oxidized to the corresponding trioxides. A mixture of antimony trisulphide and trioxide, the so-called *mineral kermes*, used in medicine, is made by boiling antimony trisulphide with sodium carbonate.

Thioarsenites and thioantimonites.—Antimony and arsenic sulphides are soluble in the alkali hydroxides, ammonium carbonate, and alkali sulphides, whereas bismuth sulphide is practically insoluble in these reagents. Similar remarks apply to the effect of these reagents upon the oxides of these elements. Curiously enough, a compound labelled "insoluble" in a given reagent often dissolves to a certain extent if another soluble compound be present, and some uncertainty is often introduced in analytical processes if this fact be ignored. Bismuth sulphide is soluble in concentrated sodium sulphide, but is reprecipitated on dilution. Unstable compounds of bismuth sulphide with the alkalis can be obtained by fusing the two substances together. With the alkali sulphides, arsenic trisulphide forms compounds called *thioarsenites*, which may be advantageously regarded as salts of an hypothetical ortho-thioarsenious acid, H_3AsS_3 , or $\text{As}(\text{SH})_3$: $\text{As}_2\text{S}_3 + 3(\text{NH}_4)_2\text{S} = 2\text{As}(\text{SNH}_4)_3$; and, as in the case of the analogous oxy-acids, salts of the hypothetical pyrothioarsenious acid, $\text{H}_4\text{As}_2\text{S}_5$; and meta-thioarsenious acid, HAS_2 , have been prepared. The thioarsenites of the alkalis, alkaline earths, and of magnesium are soluble in water, but decompose on boiling. The solutions are also decomposed by acids with the evolution of hydrogen sulphide, and the re-precipitation of arsenic trisulphide. Hydrogen sulphide will not precipitate arsenic sulphide from the normal arsenites, and it is only partially precipitated from the primary and secondary arsenites, because soluble thioarsenites are formed. If sufficient acid be present to prevent the formation of the soluble arsenites, precipitation is complete. When arsenic trisulphide is dissolved in an alkaline hydroxide, both arsenite and thioarsenite are formed: $2\text{As}_2\text{S}_3 + 4\text{KOH} = 3\text{KAsS}_2 + \text{KAsO}_2 + 2\text{H}_2\text{O}$. Similar results apply to antimony sulphide in which case thioantimonites are formed.

Arsenic disulphide, As_2S_2 , is found in nature as the mineral *realgar*; and it can be made by heating together sulphur and arsenic in the right proportions, or by heating arsenic sulphide with arsenic; or arsenious oxide and sulphur; or distilling a mixture of iron pyrites and arsenical pyrites. Arsenic disulphide is a red brittle vitreous or crystalline solid, it fuses at 307° and sublimes unchanged. Heated in air it burns with a blue flame forming sulphur dioxide and arsenic trioxide. It is used in pyrotechny; for instance, in the manufacture of the so-called "Bengal fire," which is a mixture of realgar, sulphur, and nitre.

Arsenic pentasulphide, As_2S_5 , can be made by fusing together arsenic trisulphide and sulphur in the right proportions. It is precipitated when a rapid stream of hydrogen sulphide is passed through a cold solution of arsenic acid containing a large excess of hydrochloric acid, or through a strongly acidified solution of a soluble arsenate. Both arsenic tri- and pentasulphides are reduced to the disulphide by an hydrochloric acid solution of stannous chloride: $\text{As}_2\text{S}_5 + \text{SnCl}_2 + 2\text{HCl} = \text{SnCl}_4 + \text{H}_2\text{S} + \text{As}_2\text{S}_2$. This explains the formation of the red precipitate which

occurs in the detection of arsenic in the presence of tin. The reduction is not effected by the lower chloride of arsenic, antimony, phosphorus, iron or chromium. Antimony pentasulphide, Sb_2S_5 , is made by the action of hydrogen sulphide on a mixture of antimony pentachloride with a little water. Arsenic pentasulphide, As_2S_5 , is citron yellow; antimony pentasulphide, Sb_2S_5 , is reddish yellow; and bismuth pentasulphide, Bi_2S_5 , is black. Antimony pentasulphide was once used medicinally and in veterinary work. It is used to-day in the manufacture of matches, and in vulcanizing rubber to which it imparts a red colour.

Thioarsenates and thioantimonates.—Like the corresponding trisulphides, arsenic and antimony pentasulphides dissolve in alkali sulphides, producing in the one case thioarsenates; and in the other, thioantimonates. With the alkali hydroxides, a mixture of the arsenates or antimonates with thioarsenates or thioantimonates. Crystals of ammonium thioantimonate have been made by heating antimonite with red ammonium sulphide in a sealed tube for some days and then treating the solution with alcohol. The free thio-acid is liberated by acidifying the solution of the thioarsenate or thioantimonate, but the thio-acid decomposes at once with the precipitation of the pentasulphide. Salts of ortho-thioarsenic acid—say, K_3AsS_4 —pyro-thioarsenic acid— $\text{K}_4\text{As}_2\text{S}_7$ —and meta-thioarsenic acid—say, KAsS_4 —have been reported. Similar remarks apply to the thioantimonic acids. Bismuth pentasulphide does not exhibit the acidic properties shown by the corresponding antimony and arsenic compounds. Schlippe's salt, sodium ortho-thioantimonate— $\text{Na}_3\text{SbS}_4 \cdot 9\text{H}_2\text{O}$ —is one of the best known thioantimonates. It is formed, along with sodium meta-antimonate, when antimony trisulphide is boiled with sulphur and sodium hydroxide: $4\text{Sb}_2\text{S}_3 + 8\text{S} + 18\text{NaOH} = 5\text{Na}_3\text{SbS}_4 + 3\text{NaSbO}_3 + 9\text{H}_2\text{O}$. There is a secondary consecutive reaction in which sodium thiosulphate is formed: $6\text{NaOH} + 4\text{S} = 2\text{Na}_2\text{S} + \text{Na}_2\text{S}_2\text{O}_3 + 3\text{H}_2\text{O}$; this is accentuated if the constituents be mixed and the solution boiled. The sodium thiosulphate ultimately converts the antimony sulphide into Schlippe's salt, but the time required is much greater than when the process is conducted as indicated above. Schlippe's salt crystallizes in colourless tetrahedra. The salt is decomposed by acids, depositing antimony pentasulphide. The crystals on exposure to air become covered with a yellowish-red crust of antimony pentasulphide owing to the action of the carbon dioxide in the atmosphere.

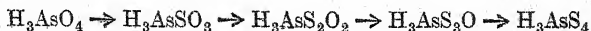
The action of hydrogen sulphide on arsenic acid.—The precipitation of pure arsenic pentasulphide from an acidified solution of arsenic acid is almost impossible because some arsenic trisulphide is formed at the same time. Concentrated hydrochloric acid in the cold favours the separation of the pentasulphide. If hydrogen sulphide be passed into a cold solution of arsenic acid, the separation of the trisulphide is very slow—over twenty-four hours are sometimes needed for the complete precipitation of the arsenic as trisulphide. The liquid remains clear for some time; then becomes turbid; and finally arsenic trisulphide, mixed with more or less pentasulphide, is precipitated. Solutions containing arsenious acid react with hydrogen sulphide at once, but with solutions containing arsenic acid, it is supposed that three consecutive reactions occur:

(1) The slow formation of a thioarsenic acid, H_3AsSO_3 , thus: $\text{H}_2\text{S} + \text{H}_3\text{AsO}_4 = \text{H}_2\text{O} + \text{H}_3\text{AsSO}_3$.

(2) The reduction of the thioarsenic acid, H_3AsSO_3 , to arsenious acid, H_3AsO_3 , with the separation of sulphur: $\text{H}_3\text{AsSO}_3 = \text{H}_3\text{AsO}_3 + \text{S}$.

(3) The decomposition of the arsenious acid by the hydrogen sulphide: $2\text{H}_3\text{AsO}_3 + 3\text{H}_2\text{S} = 6\text{H}_2\text{O} + \text{As}_2\text{S}_3$.

The initial and final stages of the reaction are represented by the equation: $2\text{H}_3\text{AsO}_4 + 5\text{H}_2\text{S} = 2\text{S} + 8\text{H}_2\text{O} + \text{As}_2\text{S}_3$. The rate of precipitation of the arsenic trisulphide is determined by the rate of reduction of the arsenic acid to arsenious acid, by the hydrogen sulphide. If the solution be heated, the decomposition of the thioarsenic acid is accelerated; but for the rapid precipitation of arsenic trisulphide it is best to reduce the arsenic acid to arsenious acid rapidly by boiling with sulphurous acid before the hydrogen sulphide is used, and not to throw the work of reduction on the slower reducing agent—hydrogen sulphide. The oxygen atoms of arsenic acid, H_3AsO_4 , are gradually replaced by sulphur atoms forming the sequence:



and the latter then dissociates into hydrogen sulphide and arsenic pentasulphide: $2\text{H}_3\text{AsS}_4 \rightleftharpoons \text{As}_2\text{S}_5 + 3\text{H}_2\text{S}$.

§ 13. The Nitrogen Family of Elements.

The properties of the group of elements—nitrogen, phosphorus, arsenic, antimony, and bismuth—just considered, show a gradual transition from non-metallic nitrogen to metallic bismuth. The apparent chemical inactivity of nitrogen is in striking contrast with the activity of phosphorus, and the volatility and insolubility of the oxides and acids of nitrogen are in equally striking contrast with the corresponding compounds of phosphorus. The identity of the empirical formulæ of these compounds does not necessarily imply that they have analogous constitutional formulæ; in fact, we are sure that the oxides are not alike in constitution. The relationship of the physical properties of the elements is best shown by arranging them in the order of their atomic weights, when the gradual and regular difference between the properties of successive elements is made clear:—

TABLE XLVI.—PROPERTIES OF THE NITROGEN-PHOSPHORUS FAMILY.

	Nitrogen.	Phosphorus.	Arsenic.	Antimony.	Bismuth.
Atomic weight .	14.01	31.0	75.0	120.2	208
Specific gravity .	1.026	1.82–2.31	4.69–5.73	6.62	9.78
Atomic volume .	13.7	about 15	13 to 16	18	21.3
Melting point .	–214°	44°	[480°]	629.2°	270°
Boiling point .	–193°	287°	[817°]	1440°	1435°

The gradual transition of non-metals into metals with increasing atomic weight is brought out very clearly in this family group. The changes in the melting and boiling points are not so regular as the other properties, but this may be related to the fact that phosphorus, arsenic, and antimony

have four-atom molecules, whereas bismuth, like nitrogen, has probably a two-atom molecule. The heat of combination of the different elements with hydrogen shows a somewhat similar gradation: NH_3 , + 12 Cals.; PH_3 , + 4.9 Cals.; AsH_3 , - 44.2 Cals.; SbH_3 , - 86.8 Cals. Ammonia is a relatively strong base, phosphine a feeble base, while arsine and stibine do not show basic qualities. Hence, as the atomic weight increases, the basicity of the hydride decreases. This is shown by the formation of the ammonium and phosphonium salts, and the non-existence of the arsenicum and stibium compounds. The resemblances between nitrogen, phosphorus, arsenic and antimony as foreshadowed by the similarities between their trihydrides is accentuated by a comparison of the compounds $\text{N}(\text{C}_2\text{H}_5)_3$, $\text{P}(\text{C}_2\text{H}_5)_3$, $\text{As}(\text{C}_2\text{H}_5)_3$, $\text{Sb}(\text{C}_2\text{H}_5)_3$, and the analogy becomes still further pronounced in comparing $\text{N}(\text{C}_2\text{H}_5)_4\text{OH}$, $\text{P}(\text{C}_2\text{H}_5)_4\text{OH}$, $\text{As}(\text{C}_2\text{H}_5)_4\text{OH}$, $\text{Sb}(\text{C}_2\text{H}_5)_4\text{OH}$, for these bodies are so surprisingly similar that they are almost indistinguishable from one another. As Hofmann and Cahours have said (1857): "They might moreover be confounded with potassium or sodium hydroxide, by which they are scarcely surpassed in alkaline power. Only the deportment of their hydrated oxides under heat enables the derivatives of phosphorus, arsenic, and antimony to be distinguished from those of nitrogen. The latter is the most stable." While phosphoric oxide is volatile and stable, the corresponding pentoxides of arsenic, antimony, and bismuth are resolved by heat into the corresponding trioxides and oxygen. Phosphorus pentoxide yields a whole series of polybasic acids, but nitrogen pentoxide gives but one monobasic acid, HNO_3 ; the corresponding salts are also in strong contrast, for the nitrates are nearly all soluble in water, while the normal, pyro-, and metaphosphates are nearly always sparingly soluble. Similarly with the trioxides, strict analogy is absent. The oxyacids diminish in strength during the passage from nitric to phosphoric, to arsenic, to antimonitic, to bismuthic acid. The acidic properties of the oxides likewise diminish from the strongly acid nitrogen and phosphorus acids to the feebly acidic oxides of arsenic, antimony, and bismuth. The basic property first appears with arsenic and becomes stronger during the passage to antimony and to bismuth. The trichloride of nitrogen is very unstable, while with bismuth the trichloride is quite stable. The boiling points of the trichlorides increase with increasing atomic weights of the elements, thus: NCl_3 boils at 71° ; PCl_3 , 78° ; AsCl_3 , 130.2° ; SbCl_3 , 200° ; and BiCl_3 , at 447° . The halogen compounds of nitrogen and phosphorus are completely hydrolyzed by water; arsenic trichloride is also hydrolyzed by water, but it can exist in solution in the presence of hydrochloric acid. Antimony and bismuth halides are incompletely hydrolyzed by water. The study of the progressive changes in the properties of these elements could be developed much further, but sufficient has been indicated in what precedes to emphasize the family resemblances.

§ 14. Vanadium, Niobium or Columbium, and Tantalum.

	Vanadium, V	Columbium, Cb	Tantalum, Ta
Atomic weights	51.06	93.5	181

The three elements—vanadium, V, Columbium, Cb, and tantalum, Ta—have many properties which relate them with the nitrogen family. The

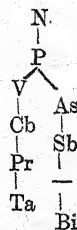
physical and the chemical properties of the three elements have a family likeness. With the physical properties :

TABLE XLVII.—PROPERTIES OF THE VANADIUM FAMILY.

	Vanadium.	Columbium, or Niobium.	Tantalum.
Atomic weight . . .	51.06	93.5	181
Colour	Silvery metal	Steel grey metal	Iron grey metal
Specific gravity . . .	5.5	7.06	8
Atomic volume . . .	9.3	13.3	23
Melting point . . .	1730°	1950°	2850°

The relations of these elements with the nitrogen family are often emphasized by the scheme shown in the margin below, where the hyphen represents a supposed element not yet discovered.

History of vanadium.—Del Rio, in 1801, expressed the opinion that a Mexican ore which he analyzed contained a new metal which he called "erythronium"—from the Greek *ἐρυθρός* (erythros), red—because it furnished red salts when treated with acids. Del Rio abandoned his opinion when Collet-Descostils, four years later, claimed that the supposed new metal was an impure chromium oxide. In 1830 N. G. Sefström described a new mineral which he found in some Swedish iron ores, and suggested for it the name "vanadium"—from "Vanadis," a Scandinavian goddess. Immediately afterwards, F. Wöhler, 1831, established the identity of Sefström's "vanadium" with Del Rio's "erythronium." F. Wöhler had also found something strange in a Mexican lead ore from Zimapan, before Sefström's announcement, and put it aside for future examination. As Berzelius expressed it :



In the far north, there lived a goddess Vanadis, as beautiful as she was gracious. One day there came a knock at the door. The goddess was in no hurry and thought, "They can knock again"; but there came no further knocking, for he who knocked had passed on. The goddess wondered who it could be who cared so little to be let in and ran to the window and recognized the departing one in the person of Wöhler. Some days afterwards some one else knocked repeatedly and loud. The goddess opened the door herself to Sefström, and vanadium was discovered.

J. J. Berzelius, 1831, investigated vanadium, and he appears to have been under the impression that the oxide VO was the metal itself. H. E. Roscoe, in 1867, isolated the metal and established its relationship with the nitrogen family of elements.

Vanadium.—Vanadium occurs in a few minerals, e.g. *vanadinite* or *vanadite*, $\text{Pb}_5(\text{VO}_4)_3\text{Cl}$, which is isomorphous with *pyromorphite*, $3\text{Pb}_3(\text{PO}_4)_2 \cdot \text{PbCl}_2$, or $\text{Pb}_5(\text{PO}_4)_3\text{Cl}$; and the corresponding arsenic compound, $3\text{Pb}_3(\text{AsO}_4)_2 \cdot \text{PbCl}_2$, or $\text{Pb}_5(\text{AsO}_4)_3\text{Cl}$; *carnotite*; *roscoelite*; *patronite*; etc.

Most of the vanadium of commerce is extracted from *vanadinite* of the Peruvian Andes which contains 8–21 per cent. of V_2O_5 . *Carnotite* is a yellow potassium uranyl vanadate. It is used for the extraction of radium and it is purchased on its radium content. Most of this ore comes from S. W. Colorado. *Roscoelite* is a vanadiferous mica; and *patronite* is

an impure vanadium sulphide. The ashes of some Peruvian coals are said to contain up to 48 per cent. of vanadic oxide, and some vanadium is extracted from those ashes. Small quantities occur in iron ores, and traces occur in many British fireclays and in granitic rocks. When the ores are fused with sodium nitrate, sodium vanadate is formed. This can be extracted with water. When the solution is treated with an ammonium salt, an orange-coloured precipitate of ammonium meta-vanadate, NH_4VO_3 , is obtained. When this salt is heated, vanadium pentoxide, more or less impure, remains. To extract vanadium from ores decomposed by acids, boil the ore with dilute sulphuric acid (1 : 1), and add ammonium persulphate to the solution. Red vanadium hydroxide is precipitated on prolonged boiling, and thus gives a fairly pure oxide on calcination. W. von Bolton (1908) made coherent rods of vanadium, niobium, and tantalum by moulding small rods of carbon, paraffin, and the pentoxide, and baking them at high temperatures. The rods were then heated by passing a moderately strong current through them *in vacuo*. The metal is also obtained by reducing the chloride with hydrogen. Vanadium pentoxide can be fused without decomposition, and it behaves in solution as a weak acid forming a series of salts: meta-, pyro-, ortho-, tetra-, and hexavanadates. These salts are isomorphous with the corresponding phosphates and resemble nitrogen in forming several oxides. Vanadic acid and its salts are reduced by hydrogen sulphide to V_2O_4 ; by sulphur dioxide, to V_2O_3 ; and by zinc and acid, to V_2O_2 . Vanadium salts of the different oxides are coloured as indicated below:

Salts of	V_2O	V_2O_2 or VO	V_2O_3	V_2O_4 or VO_2	V_2O_5
Colour	—	lavender	green	blue	orange or yellow

The basic properties of vanadium oxides become less and less pronounced as the proportion of oxygen increases. The higher oxides exhibit acidic as well as basic properties. The element also forms a series of chlorides: VCl_2 , VCl_3 , and VCl_4 . The existence of VCl_5 is doubtful. There is also a series of vanadyl compounds or oxychlorides— $(\text{VO})_2\text{Cl}$, VOCl , VOCl_2 , and VOCl_3 . Unlike nitrogen and phosphorus, vanadium is undoubtedly a metal. As in the case of phosphorus, the metal, if heated to a high temperature, burns to the pentoxide, V_2O_5 . A sample of the metal between 97 and 98 per cent. purity melted at 1720° .

Vanadium is used in making special steels because very small quantities of vanadium modify the properties of steel by increasing the hardness and malleability of the metal. The addition of a half per cent. of vanadium, for instance, raised the tensile strength of a sample of steel from $7\frac{1}{2}$ to 13 tons per square inch. The high tensile strength and elasticity of vanadium steel has led to its use for motor vehicles, etc. The so-called **ferro-vanadium** is a special alloy of iron rich in vanadium made to facilitate the alloying of definite amounts of vanadium with the molten steel. Vanadium pentoxide is used as a substitute for gold bronze; as a photographic developer, for making black aniline dyes, etc.

History of niobium and tantalum.—In 1801 C. Hatchett analyzed some chromium minerals from Connecticut, and found an earth hitherto unknown. He named the mineral "columbite," after the place of its origin, and the element was designated "columbium." A year later, 1802, A. G. Ekeberg found a new element in some Finnish minerals resembling columbite. To this he gave the name "tantalum," from "Tantalus"

of Grecian mythology, in allusion to the "tantalizing" difficulties he encountered in dissolving the mineral in acids. In 1844 H. Rose noticed two new elements in a sample of columbite from Bodenmais: one, tantalum, is similar to Ekeberg's tantalum; and the other has been called "niobium," from Niobe, the mythological daughter of Tantalus. Hatchett's "columbium" was probably a mixture of both tantalum and niobium. The term "columbium" is often applied to Rose's niobium.

Niobium, Nb, or columbium, Cb, and tantalum, Ta.—These two elements are found associated in the isomorphous minerals *tanialite*, $\text{Fe}(\text{TaO}_3)_2$; and *columbite* or *niobite*, $\text{Fe}(\text{NbO}_3)_2$. One or both of the elements occur in several rare earths—monazite, and yttritanalite—and also in tinstone, pitchblende, wolfram, and many other minerals. Columbium forms a mono-, di-, and a penta-oxide; and tantalum a di- and a penta-oxide. Columbium forms both a tri- and a penta-chloride, while tantalum forms the penta-chloride. Both elements exhibit feeble base-forming qualities, and their chief compounds are the niobates and the tantalates. The elements are produced when the chlorides are reduced in a current of hydrogen. Tantalum forms a series of complex salts with alkaline fluorides. Thus, K_2TaF_7 is formed in rhombic needles when a solution of tantalic acid in hydrofluoric acid is treated with potassium fluoride. The sparing solubility of this salt in hydrofluoric acid enables tantalum to be separated from niobium. Columbium metal melts at about 1950° , tantalum about 2300° . The metal tantalum is used as a filament in incandescent electric lamps—"tantalum lamps." These lamps resemble the ordinary carbon filament lamps, but contain a filament of tantalum wire arranged zigzag and supported on small hooks.

Didymium or rather praseodymium (Pr) and neodymium (Nd) members of the rare earth series (*q.v.*), probably belong to this family, and lie between Columbium and tantalum. Comparatively little is known about these elements.

Questions.

1. State what would be observed, and the nature of the changes which occur, in each of the following cases (equations not required): (a) when phosphorus trichloride is added to a considerable volume of water; (b) when bismuth trichloride is added to a considerable volume of water; (c) when a hot acid solution of potassium permanganate is treated with hydrogen sulphide; (d) when an alkaline chromite solution (as KCrO_2 , a salt derived from Cr_2O_3) is treated with sodium peroxide and converted into a salt derived from CrO_3 .—*Massachusetts Inst. Technology, U.S.A.*

2. Compare the hydrides of nitrogen, phosphorus, arsenic, antimony.—*Univ. Pennsylvania, U.S.A.*

3. Describe carefully the preparation of gaseous hydride of phosphorus and compare its properties with those of the corresponding hydrides of nitrogen and arsenic.—*Sheffield, Univ.*

4. What are the chief sources of white arsenic? How can this substance be converted into (a) arsenic acid; (b) Scheele's green; (c) arsenuretted hydrogen? By what property can arsenuretted hydrogen be distinguished from ammonia and from phosphine?—*London Univ.*

5. Arsenious acid is less soluble in normal hydrochloric acid than in pure water, and increase in the concentration of the acid leads to a further diminution in the solubility of the oxide until, when the concentration of the acid reaches 3.2 normal, the solubility of the arsenious acid becomes a minimum, the solubility afterwards rising as the concentration of the acid is still further increased. Discuss and explain these results.—*St. Andrews Univ.*

6. Discuss the following changes and suggest an explanation applicable to

both of them :—(a) when concentrated hydrochloric acid is mixed with concentrated brine a precipitate is formed which disappears on the addition of water ; (b) when water is mixed with bismuth chloride solution, a precipitate is formed which disappears on the addition of concentrated hydrochloric acid.—*London Univ.*

7. Which of the elements are to be regarded as most similar to nitrogen ? Indicate the grounds on which your answer is based.—*Aberdeen Univ.*

8. Why are phosphoric and arsenic acids regarded as "tribasic acids" when they yield salts such as those of the formulae Na_2HPO_4 and Na_2HAsO_4 , on neutralization with sodium carbonate ? Formulate meta-, ortho-, and pyrophosphoric acids in accordance with (a) the old dualistic theory of the composition of acids and salts ; and (b) the modern theory. What facts apparently justify the latter ? —*London Univ.*

9. In what forms does arsenious oxide exist ? What happens when the oxide is (a) introduced into an apparatus in which hydrogen is being generated ; (b) heated ; (c) heated with charcoal ; (d) heated with dilute nitric acid ? —*London Univ.*

10. Describe the action of water (a) in small and (b) in large quantity on the chlorides of phosphorus, arsenic, and antimony.—*Board of Educ.*

11. How is arsenic obtained ; and how would you prepare arsenic chloride, arsenic trioxide and arsenic acid from arsenic ? What are the more characteristic properties of these substances ? —*Board of Educ.*

12. A white salt is found to dissolve readily in water. The usual reagents produce no change in the solution, which has a slightly alkaline reaction. It imparts a soft lilac colour to a Bunsen's flame, fusing without other visible change. With platinic chloride it gives a yellow, quickly subsiding, precipitate. With barium chloride the solution gives an abundant white precipitate, but only when rather concentrated. With silver nitrate it gives a precipitate of a pink brown colour. After adding sulphurous acid and boiling for some minutes, on testing with hydrogen sulphide there is a bright yellow flocculent precipitate. What is the salt ? —*New Zealand Univ.*

13. What is the evidence for the existence of antimonyl and bismuthyl compounds ? Describe the preparation of an antimonyl compound from antimony, and of a bismuthyl compound from bismuth ; give briefly the properties of each ; and explain the reactions that occur when each is boiled with sodium carbonate solution.—*Punjab Univ.*

14. Why are the vapour densities of phosphorus, arsenic, cadmium, and mercury considered to be anomalous ? The molecular weight of nitric oxide being taken as 30, what is the density of the gas compared with that of hydrogen ? —*London Univ.*

15. Describe Marsh's test for arsenic. If a mixture of arsenuretted and antimonuretted hydrogen be passed into a solution of silver nitrate, decompositions occur which enable the arsenic and antimony to be separated and distinguished. Explain the reactions involved.—*London Univ.*

CHAPTER XXXIII

BORON, ALUMINIUM, AND RELATED ELEMENTS

§ 1. Boric Acids and the Borates.

Extraction.—Volcanic jets of steam (*soffioni*) at a temperature between 90° and 120° , issuing from the fumaroles on the so-called Maremma di Toscana—or Tuscany Marshes—carry small quantities of boric acid. The steam condenses in lagoons (*lagoni*) of water which often surround the jets. The water of the lagoons becomes highly charged with the acid, and the boric acid can be obtained in a crystalline condition by evaporating the water of the lagoons. Artificial lagoons for arresting the jets of steam were established in Tuscany in 1818; and artificial *soffioni* were bored in 1854. The natural heat of the steam is utilized in concentrating the water. The crude acid—called Tuscany acid—is contaminated with ammoniacal salts, etc., and it is coloured yellowish brown. The crude acid contains about 82.5 per cent. of boric acid, and it can be refined by dissolving it in hot water, and mixing the solution with freshly ignited powdered charcoal. This removes the colouring matters. The clear solution is then allowed to crystallize, and this is placed on the market as “refined boric acid” or “refined boracic acid.” The water of the hot springs at Sulphur Bank and other places in Nevada and California is rich in borates.

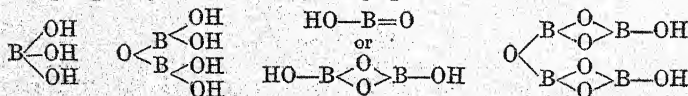
Action of solvents.—Boric acid crystallizes during the cooling of its hot aqueous solutions as white shining plates belonging to the triclinic system. When the aqueous solution is boiled, the acid volatilizes with the steam; 100 grams of water at 100° dissolve 27.5 grams of H_3BO_3 ; at 50° , 8.8 grams; and at 0° , 1.95 grams. Boric acid is more soluble in alcohol than in water. The alcoholic solution burns with a characteristic flame tinged with green, and in this way, 0.001 gram of boric acid can be detected in a solution containing 0.1 per cent. of the acid. The aqueous solution turns blue litmus claret-red; and moist yellow turmeric paper is coloured brown. Unlike the brown colour produced by alkalis with turmeric paper, the boric acid stain is not destroyed by acids. In a solution acidified with hydrochloric acid, 0.0001 gram of boric acid in a solution containing 0.01 per cent., will give a pink coloration when the paper is dried at 100° ; if too much acid be present, the stain will be brown, and if no acid be present, the colour does not show.

Titration of boric acid.—As is the case with carbonic acid, boric acid does not effect the colour of methyl orange. If phenolphthalein be used as indicator, and the boric acid be titrated with sodium hydroxide, the pink colour of the phenolphthalein appears before all the boric acid is neutralized. This

is due to the hydrolysis of the alkaline borate. If mannite or glycerol (free from acid) is added to the solution, the hydrolysis is prevented, hence the end point of the titration is reached when the red colour of the phenolphthalein persists when more mannite is added. The reaction which occurs during the titration is: $\text{NaOH} + \text{H}_3\text{BO}_3 = \text{NaBO}_2 + 2\text{H}_2\text{O}$.

Action of heat.—Boric acid is a tribasic acid, and it has the ultimate composition H_3BO_3 , or $\text{B}(\text{OH})_3$. A solution containing 22.38 per cent. raised the boiling point of water 1.90° ; hence, from the formula on p. 256, $5.2 \times 22.38 \div 1.90 = 62.7$. The theoretical number for H_3BO_3 is 62. Hence with boron trivalent, this acid is orthoboric acid, $\text{B}(\text{OH})_3$. If heated to 107° – 108° , orthoboric acid loses water and forms metaboric acid, HBO_2 , also written $\text{H}_2\text{B}_2\text{O}_4$. 1.1 gram of metaboric acid dissolved in water lowered the freezing point 0.490° . Hence from the formula on p. 259, $18.5 \times 1.1 \div 0.49 = 42.8$, corresponding with the formula HBO_2 . If heated to 138° – 140° , boric acid forms tetraboric acid, or the so-called "pyroboric acid," $\text{H}_2\text{B}_4\text{O}_7$: $4\text{H}_3\text{BO}_3 = 5\text{H}_2\text{O} + \text{H}_2\text{B}_4\text{O}_7$. When boric acid is heated to a still higher temperature—"red heat"—it forms anhydrous boron trioxide, or boric oxide. The fused mass solidifies to a colourless transparent glass which absorbs moisture from the atmosphere, and becomes opaque, finally passing into boric acid. (Anhydrous silica does not combine directly with water.) Boric oxide volatilizes extremely slowly at a red heat. Sulphuric acid, it will be remembered, is not so strong an acid as hydrochloric acid at ordinary temperatures, but, owing to the more volatile nature of hydrochloric acid, the less volatile and weaker sulphuric acid is able to convert the chlorides into sulphates, as indicated previously. Similarly, in consequence of the non-volatility of boric acid at a red heat, it is able to decompose the compounds of the stronger acids with the bases and form borates: $\text{B}_2\text{O}_3 + 3\text{K}_2\text{SO}_4 = 2\text{B}(\text{OK})_3 + 3\text{SO}_3$. Sulphuric and hydrochloric acids are much stronger than boric acid when competing for the bases in aqueous solutions, but at a red heat, the more volatile acids are driven off.

Constitution.—The three boric acids are related by the empirical formulae: orthoboric acid, $\text{B}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$; pyroboric acid, $\text{B}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$, metaboric acid, $\text{B}_2\text{O}_3 \cdot \text{H}_2\text{O}$; and tetraboric acid, $2\text{B}_2\text{O}_3 \cdot \text{H}_2\text{O}$. The relationship is perhaps best shown by the graphic formulae:



Orthoboric acid, $\text{B}(\text{OH})_3$, Pyroboric acid, $\text{H}_2\text{B}_4\text{O}_7$, Metaboric acid, $\text{BO}(\text{OH})_2$, or $\text{H}_2\text{B}_2\text{O}_4$.

Tetraboric acid, $\text{H}_2\text{B}_4\text{O}_7$.

The meta- and pyro-boric acids pass to orthoboric acid in aqueous solution. While derivatives of the hydrates are known, only the ortho-acid can exist in solution for any length of time.

Orthoborates.—The salts of orthoboric acid are unstable and few well-defined orthoborates are known. Magnesium orthoborate, $\text{Mg}_3(\text{BO}_3)_2$, is one of the best known orthoborates. There is no doubt about the tribasicity of the acid, however, because the hydrogen of $\text{B}(\text{OH})_3$ has been successively replaced by organic radicles such as C_2H_5 . Thus, ethyl orthoborate has a vapour density corresponding with $\text{B}(\text{OC}_2\text{H}_5)_3$. This latter compound is formed as a volatile compound in the flame test

for boric acid. The mineral *ascharite*, MgHBO_3 , is magnesium hydrorthoborate.

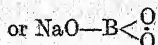
Metaborates.—The metaborates appear to be the most stable salts of boric acid. Potassium forms a metaborate, KBO_2 ; and a series of potassium acid borates— $\text{KBO}_2 \cdot \text{HBO}_2$, or KHB_2O_4 ; and the $\text{KBO}_2 \cdot 3\text{HBO}_2$.

Tetraborates or pyroborates.—These are also stable compounds. The best known is borax, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$. A series of complex calcium and magnesium borates occur in nature: *borocalcite*, $\text{CaB}_4\text{O}_7 \cdot 6\text{H}_2\text{O}$, from Iquique; *boronatrocalcite* (ulexite), $\text{NaCaB}_4\text{O}_9 \cdot 8\text{H}_2\text{O}$, from Chili; *boracite*, $2\text{Mg}_3\text{B}_5\text{O}_{15} \cdot \text{MgCl}_2$, from the Stassfurt deposits; *pandermite*, $\text{Ca}_5\text{B}_{20}\text{O}_{38} \cdot 15\text{H}_2\text{O}$, from Panderna in Asia Minor; *colemanite* $\text{Ca}_2\text{B}_6\text{O}_{11} \cdot 5\text{H}_2\text{O}$; etc.

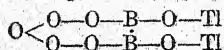
The borates nearly always contain water of crystallization, and, excepting the alkaline borates, they are but slightly soluble in water. The borates are easily decomposed by water, and acidulated solutions of the borates give the flame and other reactions characteristic of boric acid. In some places boric acid is made directly from the native calcium borates by dissolving the mineral in hot hydrochloric acid and collecting the crystals of boric acid which separate on cooling. The acid can also be made by the action of sulphuric or hydrochloric acid on borax.

Boric oxide, like alumina, is an "intermediate oxide," for it has feeble acidic and feeble basic properties. As a weak acid it forms a series of salts—borates. Boric oxide also forms complex salts—boro-silicates—with the metallic oxides and silica, e.g. *tourmaline*. Boric oxide probably plays the same rôle in the boro-silicates as it plays in the boro-tungstates, and alumina in the alumino-silicates. Corresponding with the basic properties of the oxide, unstable compounds like the acid sulphate, $\text{B}(\text{HSO}_4)_3$; phosphate, BPO_4 , etc., have been formed. Most of the salts are hydrolyzed by water, dilute acids, and dilute alkaline solutions.

Perborates.—When a mixed solution of borax and sodium hydroxide is electrolyzed, or treated with hydrogen peroxide, sodium perborate $\text{NaBO}_3 \cdot 4\text{H}_2\text{O}$ is formed—possibly $\text{NaO}-\text{O}-\text{B}=\text{O}$



related to sodium metaborate $\text{NaO}-\text{B}=\text{O}$. The perborates act as strong oxidizing agents. Another form of perborate is illustrated by thalious perborate, $\text{Th}_2\text{B}_4\text{O}_7$, formed by the action of hydrogen peroxide on a mixture of thallic sulphate and boric acid. It decomposes with the evolution of oxygen on warming. The graphic formula is possibly



The bleaching action of perborates is not attended by the deleterious effects on the fabrics attended by most other bleaching agents. The perborates are far less liable to deteriorate on storing than the peroxides, and hence the perborates are superior for many bleaching and disinfecting purposes.

§ 2. Borax.

Manufacture.—Borax occurs as a natural deposit—*tincol*—in the dried-up inland lakes in some parts of India, Tibet, and California—e.g. the

"borax marsh" in the Saline Valley. Native tincal contains about 55 per cent. of borax proper: $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$. This is extracted by lixiviating the mass with water, and evaporating the clear solution until the crystals separate. Before 1870, most of the borax used in Europe was made by digesting a solution of boric acid with the proper amount of soda ash, whereas the borax now in commerce is usually made from native calcium borates. The higher grade ores run to the equivalent of 45 per cent. crystalline borax. The powdered mineral is boiled for hours in drums with a slight excess of an aqueous solution of sodium carbonate. Calcium carbonate is precipitated as a "mud": $\text{Ca}_2\text{B}_6\text{O}_{11} + 2\text{Na}_2\text{CO}_3 \rightleftharpoons 2\text{CaCO}_3 + \text{Na}_2\text{B}_4\text{O}_7 + 2\text{NaBO}_2$. On crystallizing the clear solution, crystals of borax are obtained. Carbon dioxide is blown through the mother liquid to convert the sodium metaborate into borax: $4\text{NaBO}_2 + \text{CO}_2 \rightleftharpoons \text{Na}_2\text{CO}_3 + \text{Na}_2\text{B}_4\text{O}_7$. The borax separates in fine crystals, leaving sodium carbonate in the solution. The residual "mud" is compressed in a filter press to remove the soluble matters it contains.

The action of water.—100 grams of water at 10° dissolve 1.6 grams of borax, $\text{Na}_2\text{B}_4\text{O}_7$; at 50° , 10.5 grams; and at 100° , 52.5 grams. The aqueous solution has a feebly alkaline reaction owing to hydrolysis. In concentrated aqueous solutions: $\text{Na}_2\text{B}_4\text{O}_7 + 3\text{H}_2\text{O} \rightleftharpoons 2\text{NaBO}_2 + 2\text{H}_3\text{BO}_3$; and on further dilution, hydrolysis may be practically completed: $\text{NaBO}_2 + 2\text{H}_2\text{O} \rightleftharpoons \text{NaOH} + \text{H}_3\text{BO}_3$. If a saturated solution be allowed to crystallize above about 62° , octahedral crystals of the pentahydrate, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 5\text{H}_2\text{O}$, separate; and if the temperature be below this transition point ordinary monoclinic crystals of the decahydrate, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$, are obtained.

Titration of borax.—A dilute solution of borax may be titrated with standard hydrochloric acid until all the soda is neutralized and the boric acid is all free, provided methyl orange be used as indicator. The reaction is: $\text{Na}_2\text{B}_4\text{O}_7 + 2\text{HCl} + 5\text{H}_2\text{O} = 2\text{NaCl} + 4\text{H}_3\text{BO}_3$. The boric acid has no effect on the methyl orange. If the same amount of the standard acid be added to the same amount of the borax solution, and phenolphthalein be used as indicator, the freed boric acid can be titrated as indicated previously. In this manner, the amount of alkali, and the amount of boric acid in a sample of borax can be determined.

The action of heat.—When heated, borax fuses, loses water, and swells up into a white porous mass, owing to the expulsion of the water. Finally the borax melts to a clear glass—borax glass—which is anhydrous borax. Like boric acid, fused borax dissolves many colouring oxides, giving glasses with a characteristic colour. Borax "beads" fused in a loop of platinum wire are used as tests for oxides which dissolve in the bead, and show characteristic colours as indicated in Table. Thus:

TABLE XLVIII.—COLOURS OF BORAX BEADS WITH SOME METALLIC OXIDES.

Metallic oxide.	Oxidizing flame.	Reducing flame.
Copper	Green (hot); blue (cold)	Colourless or red.
Cobalt	Blue (hot or cold)	Blue.
Chromium	Green or red (hot or cold)	Green.
Iron	Yellow (cold); brown (hot)	Dirty green or olive.
Nickel	Violet (hot); brown (cold)	Grey and opaque.
Manganese	Amethyst (hot or cold)	Grey and opaque.

The dissolution of metallic oxides in fused borax is supposed to be connected with the fact that borax contains an excess of acid anhydride: $2\text{NaBO}_3 \cdot \text{B}_2\text{O}_3$. The union of the metallic oxide with the boric oxide forms a metaborate. Thus, with copper oxide, $\text{Na}_2\text{B}_4\text{O}_7 + \text{CuO} = \text{Cu}(\text{BO}_2)_2 + 2\text{NaBO}_2$.

Uses.—Large quantities of borax are used in the manufacture of enamels, glazes, and of optical glass; in preserving wood and articles of food if its use is not restricted by law; as a mild antiseptic for making lotions and ointments; in the manufacture of soap; as a cleansing and stiffening agent in laundry work; for stiffening candle wicks; in making varnishes for metals; manufacture of drying oils; with casein as a substitute for gum arabic; as a flux; and in soldering and brazing. Solder only adheres to clean metal surfaces, and molten borax will often dissolve the contaminating oxides which prevent solder adhering.

§ 3. Boron.

Isolation of the element.—The element boron is obtained by the electrolysis of fused borax with iron electrodes, and by heating the oxide with sodium, or with potassium, magnesium, or aluminium in a covered crucible: $\text{B}_2\text{O}_3 + 6\text{K} = 3\text{K}_2\text{O} + 2\text{B}$. The fused mass is boiled with dilute hydrochloric acid, and a dark brown powder of amorphous boron remains. By using magnesium powder, H. Moissan (1892) made a sample of amorphous boron containing 99·6 per cent. of the element, and but 0·4 per cent. of impurity. Crystalline boron can also be made as well as the amorphous powder. By dissolving boron in molten aluminium at a high temperature, the solution, on cooling, deposits crystals of boron. The aluminium can be removed by boiling the mass with a solution of sodium hydroxide. The insoluble boron remains behind in transparent yellow, or yellowish-brown (monoclinic) crystals. Crystalline boron is always contaminated with a little aluminium and carbon. Some think that the alleged crystalline boron is really a compound of aluminium and boron— AlB_{12} , or $\text{B}_{12}\text{C}_2\text{Al}_3$, for example. It is insoluble in most solvents.

Properties.—Boron melts at about 2200° , and it volatilizes a little at that temperature. It is oxidized by carbon monoxide and by silica. It is attacked by the oxy-acids, and by steam. It is said to be slightly soluble in water, giving the solution a brown colour from which it is precipitated by acids and salts. These facts make it probable that the solution is colloidal (p. 298). Crystalline boron is not attacked by concentrated nitric acid; and aqua regia attacks it but slowly. Boiling sodium hydroxide has no appreciable action, but fused sodium hydroxide dissolves it slowly, forming sodium borate and hydrogen.

Compounds with other elements.—Boron burns to the trioxide when heated to about 700° in air, and when heated in nitrogen or ammonia it forms boron nitride, BN. This compound is decomposed by steam: $\text{BN} + 3\text{H}_2\text{O} = \text{B}(\text{OH})_3 + \text{NH}_3$. The occurrence of ammonia and boric acid in the vapours of the soffioni of Tuscany has led to the theory that these compounds are derived from the action of steam on subterranean boron nitride. Boron combines with carbon and silicon at the temperature of the electric arc furnace. Carbon boride, CB_2 , and silicon

borides, SiB_3 , and SiB_6 , are very hard crystalline substances, which in many ways resemble the diamond and crystalline boron and silicon. They resist attack by most chemical agents. Metallic borides are formed by the direct union of the elements—Pt, Ag, etc.—at a high temperature. Thus, magnesium boride, Mg_3B_2 , more or less impure, is made by strongly heating a mixture of boron trioxide with magnesium powder: $6\text{Mg} + \text{B}_2\text{O}_3 = 3\text{MgO} + \text{Mg}_3\text{B}_2$. Calcium boride in contact with water gives hydrogen and hydrogen boride, and thus behaves unlike calcium carbide which gives acetylene. The alloys of iron and boron—iron borides—unlike the ferrosilicons are scarcely attacked by hydrochloric or hydrofluoric acid. Boron also combines with sulphur, forming boron sulphide, B_2S_3 . This compound is best made by passing the vapour of carbon disulphide over a mixture of carbon and boron trioxide at a red heat: $2\text{B}_2\text{O}_3 + 3\text{C} + 3\text{CS}_2 = 6\text{CO} + 2\text{B}_2\text{S}_3$; by the action of sulphur vapour on amorphous boron; by the action of hydrogen sulphide on boron at a red heat; and by heating boron iodide with sulphur; or, best of all, by heating iron boride in a stream of hydrogen sulphide. The free sulphur can be removed by carbon disulphide. Boron sulphide is a white crystalline solid with an unpleasant smell. The vapour attacks the eyes. In contact with water, the sulphide is immediately hydrolyzed, forming boric acid and hydrogen sulphide: $\text{B}_2\text{S}_3 + 6\text{H}_2\text{O} = 2\text{B}(\text{OH})_3 + 3\text{H}_2\text{S}$. Carbon dioxide decomposes it: $\text{B}_2\text{S}_3 + 3\text{CO}_2 = \text{B}_2\text{O}_3 + 3\text{CO} + 3\text{S}$. Boron pentasulphide is prepared by heating on a water-bath an excess of a solution of sulphur in carbon disulphide with one of boron iodide in the same solvent. The white crystals melt at 390° . Boron phosphide, PB , is made from the elements. It takes fire in cold chlorine gas, and also when treated with concentrated nitric acid; when heated in a current of hydrogen at 100° , it passes into the more stable P_3B_5 , which is not attacked in the cold by chlorine or nitric acid. For compounds with the halogens, see below.

Occurrence.—The element does not occur free in nature, but it occurs as boric acid, also called Tuscan boric acid; tincal; boracite; and the complex calcium borates, already discussed, are the chief modes in which boron occurs. Boric oxide has been reported in sea-water—0.2 gram per cubic metre. It is also found in small quantities in soils; in most vegetable products; in plant ashes; and in some wines. It is not so common in animal products. H. Jay could detect none in blood, but he found it in horn. It also occurs in some spring-waters—e.g. the Steamboat Springs, Nevada.

Atomic weight.—The combining weight of boron has been determined from the weight of water present in a given weight of crystalline borax, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$; and by the analysis of boron carbide, chloride, bromide, and sulphide, etc. The results show that if oxygen be taken as 16, boron has a combining weight ranging between 10.82 and 11.05; the best representative value is supposed to be 11. Dulong and Petit's rule cannot be used because of the "abnormal" specific heat of this element, Fig. 233. Several volatile compounds of boron are known, and the vapour densities of a number of these compounds lead to the inference that the number 11 also represents the atomic weight.

History.—Although mentioned in the early Latin writings on chemistry, it is probable that the term "borax" did not always refer to the substance

now called "borax," since the Arabians applied the term *baurach* (borax) indifferently to borax and to nitre. In 1702 W. Homberg made boric acid from borax, and called the acid *sal sedativum*; J. H. Pott (1741) showed that ordinary Glauber's salt was produced at the same time. Hence, sulphuric acid and borax yields boric acid and Glauber's salt. Baron, 1748, showed that borax is a compound of Homberg's *sal sedativum* and soda. After Lavoisier's work on acids, the term "boracic acid" was substituted for "*sal sedativum*," and "boracic acid" was later abbreviated to "boric acid." J. L. Gay-Lussac and J. Thénard isolated the element in a more or less impure condition in 1808.

§ 4. Boron Hydrides, and Halides.

Boron hydrides.—When magnesium boride is slowly added to dilute hydrochloric acid at about 50° while a stream of hydrogen passes through the apparatus, the gases which are evolved, when cooled by liquid air, furnish a white crystalline solid with the empirical formula BH . According to A. Stock (1913–15), this solid is a mixture of various boron hydrides, with silicon hydride, carbon dioxide, and other impurities. By fractional distillation of the product, tetraboron, hydride B_4H_{10} , boiling at 16° – 17° under ordinary pressure, and a hexaboron hydride, B_6H_{12} , boiling at about 100° were obtained. The former decomposes at 100° , forming diboron hexahydride, a gas boiling between -87° and -88° . If the diboron hydride is heated at 115° – 120° for some time it forms decaboron hydride $B_{10}H_{14}$, a white solid melting at 99.5° . Three other hydrides have been obtained, but nothing corresponding with the *triborene*, B_3H_3 , or with the BH_3 or B_2H_4 were noticed. The formula B_4H_{10} has been established from the following facts:—(i) it rapidly decomposes in contact with water, furnishing products corresponding with the equation: $B_4H_{10} + 12H_2O \rightarrow 4H_3BO_3 + 11H_2$; (ii) by explosion with an excess of air; and (iii) the vapour density is 55.6 (theory for $B_4H_{10} = 56$). Hence, boron must be at least quadrivalent, and not trivalent as a maximum; the graphic formula being $H_3 \equiv B - B \equiv H_3$. The constitution of B_4H_{10} on the assumption that boron is quadrivalent will be represented by $BH_3.BH_2.BH_2.BH_3$; and that of B_6H_{12} , by $BH_3.BH_2.BH_2.BH_2.BH_2.BH_3$. For some properties of the boron hydrides—*vide* silicon hydrides.

Boron trihalides.—All four halogens unite with boron to form trihalides:

	BF_3	BCl_3	BBr_3	BI_3
Condition . . .	Colourless gas	Colourless liquid	Colourless liquid	White solid
Specific gravity	1.35 (17°)	2.69	3.3 (50°)
Boiling point . . .	-101° (melts – 127°)	18.2°	90.5°	210° (melts 43°)

The first three compounds can be made by the direct union of the elements; and all four can be made by heating an intimate mixture of boron trioxide and charcoal in a current of the respective gas: $B_2O_3 + 3C + 3Cl_2 = 3CO + 2BCl_3$. Boron chloride and fluoride can also be made by the respective action of hydrochloric or hydrofluoric acid upon boron, or on boric oxide; say, by heating a mixture of concentrated sulphuric acid, calcium fluoride, and boric oxide: $B_2O_3 + 6HF = 2BF_3 + 3H_2O$. A similar process with boron furnished boron trifluoride, BF_3 , but silicon tetrachloride, $SiCl_4$, was not obtained in this way. This is probably due

to the fact that boron trichloride is not quite so readily hydrolyzed as silicon tetrachloride by the water formed in the reaction. Boron trichloride boils at 18.2° , and it can be distilled over sodium without decomposition; but with magnesium, the vapour forms magnesium boride; Mg_3B_2 ; while zinc dust does not react below 200° . All the boron trihalides are hydrolyzed by water: $\text{BF}_3 + 3\text{H}_2\text{O} = \text{B}(\text{OH})_3 + 3\text{HF}$. The hydrogen fluoride, so produced, unites with some of the unchanged boron trifluoride, BF_3 , producing a complex monobasic acid—hydrofluoboric acid, HBF_4 , thus: $\text{BF}_3 + \text{HF} = \text{HBF}_4$. A series of corresponding salts called borofluorides or fluoborates have been obtained. Potassium fluoborate is dimorphous, crystallizing from hot aqueous solutions in cubic (octahedral) crystals, and from cold solutions in rhombic plates. Corresponding chloro- and bromo-borates have been reported. Boron fluoride unites with ammonia forming a series of compounds: $\text{BF}_3 \cdot \text{NH}_3$; $\text{BF}_3 \cdot 2\text{NH}_3$; $\text{BF}_3 \cdot 3\text{NH}_3$; similar compounds are obtained by the union of BF_3 with PH_3 ; of boron bromide with the phosphorus bromides, and of boron chloride with phosphorus pentachloride. An excess of water is needed for the complete hydrolysis of the trichloride. If but a small quantity of water be used, a solid hydrate is produced.

§ 5. Aluminium—Occurrence, Preparation, and Properties.

Occurrence.—Aluminium, once also called aluminum, does not occur free in nature, but its compounds are numerous and widely distributed. It comes third, after oxygen and silicon, on Clarke's list. *Corundum*, *ruby*, and *sapphire* are more or less impure forms of the oxide, Al_2O_3 ; *emery* is a mixture of iron oxide and corundum. There are three recognized hydrates occurring in nature—*gibbsite* or *hydrargillite*, $\text{Al}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$, or $\text{Al}(\text{OH})_3$; *bauxite*, $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$, or $\text{Al}_2\text{O}(\text{OH})_2$, and *diaspore*, $\text{Al}_2\text{O}_3 \cdot \text{H}_2\text{O}$, or $\text{AlO}(\text{OH})$. The word *bauxite* is used in rather a general way for native aluminium hydroxides containing the equivalent of, say, 50 to 70 per cent. of Al_2O_3 , about 25 to 30 per cent. of water, with varying amounts ferric oxide, titanic oxide, and silica. Some bauxites approximate to the trihydrate; others appear to be mixtures of the trihydrate and lower hydrates, and some are considered to be more or less impure monohydrate. The bauxites are often classed as ferruginous, and non-ferruginous. The double fluoride—*cryolite*, $\text{AlF}_3 \cdot 3\text{NaF}$, is discussed on p. 328; *turquoise* is a hydrated phosphate. The *felspars* and *clays*, as well as a great number of common minerals, are complex silicates of aluminium with other bases. Bauxite and cryolite are the chief compounds used in the preparation of the metal aluminium.

History.—The word "alumen," or its Greek equivalent, was formerly applied as a grouping term for substances with an astringent taste. Geber and some others classed alum with the "vitriols," but Paracelsus considered it to be radically different, for he pointed out that its "corpus" is not metallic but an intimate mixture of earths. The earthy "corpus" was confused with lime until J. H. Pott, 1746, showed that the base is really an argillaceous earth; and in 1754, A. S. Marggraf proved clearly that the base is entirely different from lime, and that clay contains the "alum earth" united with silica.

Isolation of the metal.—After Davy's work on potassium, calcium,

etc., many attempts were made to isolate the unknown element in alumina, and F. Wöhler succeeded in reducing anhydrous aluminium chloride, AlCl_3 , to the metal by warming the chloride with potassium: $\text{AlCl}_3 + 3\text{K} = 3\text{KCl} + \text{Al}$. Wöhler's first experiments furnished a light grey metallic powder, and it was not until 1845 that the compact metal was obtained. The metal was at that time a "chemical curiosity." A specimen exhibited at the Palais de l'Industrie in 1855 was labelled *L'argent de l'argile*. In 1855 aluminium sold at the rate of about £120 per kilogram, and in 1910 about £80 per ton. H. St. C. Deville, in 1854, reduced aluminium chloride in quantity by means of sodium in place of Wöhler's potassium; H. Rose, 1855, recommended the reduction of cryolite by sodium; and N. Beketoff, the reduction of cryolite by magnesium. These chemical processes are not now used on an industrial scale.

R. Bunsen isolated aluminium by the electrolysis of the fused chloride in 1854, and the result was confirmed by Deville in the same year. A. and E. H. Cowles Bros., 1884, reduced alumina by heating it with carbon in an electric arc furnace. The formation of aluminium carbide, Al_4C_3 , introduced some practical difficulties. The method was for a time more or less successfully employed in making alloys of aluminium with copper. The electric furnace was charged with corundum, carbon, and metallic copper. After the charge had been heated about five hours, a copper-aluminium alloy was obtained. The cheap production of aluminium was made possible by the discovery, by C. M. Hall, in 1886, that a solution of alumina in a molten mixture of cryolite and some other fusible fluoride, say, potassium fluoride, is an electrolyte, and, when electrolyzed, aluminium collects at the cathode, oxygen at the anode. Patents by C. S. Bradley (1883-1891), P. L. T. Héroult (1887 et seq.), and many others embodied ideas more or less valuable, and, as a result, aluminium can now be made cheaply on a large scale. Electrolytic processes are practically the only methods used for the preparation of aluminium.

Héroult's electrolytic process for aluminium.—In this process the electrolyte is a solution of bauxite in fused cryolite. The "electrolytic cell" is a rectangular iron box (about 8 feet long and 6 feet wide). This cell—Fig. 246—is connected with the electric generator so that the iron box itself serves as the cathode. The anode, or rather anodes, is a set of about 48 carbon rods (3 inches diameter and 15 inches long) arranged in three or four rows and suspended from copper rods connected with the other pole of the generator. The resistance of the electrolyte to the current gives enough heat to keep the mass fluid. The liquid metal sinks to the bottom of the cell, whence it is "tapped" from time to time. The oxygen evolved at the anode either escapes as a gas or unites with the carbon to form carbon monoxide which either burns or escapes. The process is continuous for fresh supplies of bauxite are added when needed. The resistance of the bath increases when the alumina wants replenishing; and this causes a lamp, shunted off the main circuit, to glow. This tells the workmen that a fresh charge of bauxite is needed.

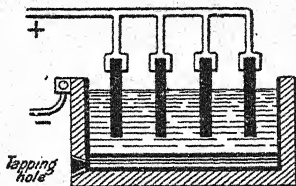


FIG. 246.—Héroult's Aluminium Furnace (Diagrammatic).

Although clays usually contain the equivalent of 20 to 36 per cent. of alumina, no method is known for separating silica from the alumina cheaply enough for use on a manufacturing scale. It is necessary to use a fairly pure bauxite for the process, otherwise the metallic aluminium will be seriously contaminated with impurities. Apart from the cost of the electrical energy, the purification of the raw bauxite is the most costly item in the process. Native bauxite is usually too impure to be used without a preliminary purification to remove the iron, titanium, and silicon oxides.

Purification of bauxite.—The bauxite may be fused with sodium hydroxide or sodium carbonate: $\text{Al}_2\text{O}_3 + 3\text{Na}_2\text{CO}_3 \rightarrow 2\text{Al}(\text{ONa})_3 + 3\text{CO}_2$, and the cold mass extracted with water; or better, the bauxite may be roasted at a low temperature to convert the ferrous oxide into ferric oxide, and then digested with a solution of sodium hydroxide in iron kettles under pressure: $\text{Al}(\text{OH})_3 + 3\text{NaOH} = \text{Al}(\text{ONa})_3 + 3\text{H}_2\text{O}$. The alumina is precipitated from the solution by *C. J. Baeyer's process*, 1887, namely by agitation with a little aluminium hydroxide, when the aluminium hydroxide separates from the solution until the ratio $\text{Al}_2\text{O}_3 : \text{Na}_2\text{O}$ is about 1:6. The aluminium hydroxide can also be precipitated by passing carbon dioxide through the solution: $2\text{Al}(\text{ONa})_3 + 3\text{CO}_2 + 3\text{H}_2\text{O} = 2\text{Al}(\text{OH})_3 + 3\text{Na}_2\text{CO}_3$. The aluminium hydroxide in either case is washed, dried, and ignited. In *Serpel's process*, by the action of water on aluminium nitride, the ammonia can be regarded as a by-product in the manufacture of alumina, for, instead of agitating bauxite with sodium aluminate, aluminium nitride can be used. The "combined iron" of the bauxite is reduced to metallic iron when the nitride is formed, and this settles as a mud when the nitride is decomposed in the solution of sodium aluminate.

Properties.—Aluminium is a bluish white metal capable of taking a high polish. The dull surface usually seen on the metal is an effect of a superficial film of oxide. Aluminium is lighter than most metals, its specific gravity is 2.6, and therefore it has nearly the same specific gravity as glass, and one-third the specific gravity of iron. The metal is ductile and malleable at 100° to 150°. With frequent annealing it can be rolled into sheets, wire, and foil. Aluminium is not very tenacious; at about 530° it undergoes a remarkable change in its physical properties, whereby its toughness and hardness are considerably reduced, and the metal becomes so friable that it can be pulverized. It is a good conductor of heat and electricity. Aluminium melts at about 657°: and boils at 1800°.

Aluminium remains practically unaltered in dry air, while in moist air and in boiling water, a superficial film of oxide seems to protect the metal from further action. Even at 700° to 800° it oxidizes but slowly; at higher temperatures it burns brilliantly with the evolution of much heat.

If a clean strip of aluminium be placed in a bottle containing a few c.c. of mercury and all shaken, the aluminium, when exposed to air, rapidly oxidizes, and white tufts of alumina grow up to about a cm. above the surface of the metal. If the aluminium be placed in water instead of in air, the water is decomposed, and the aluminium oxidized. If a little mercury be poured into a beaker of water and a clean strip of aluminium be dipped in the vessel, no action occurs until the aluminium touches the mercury. The water then decomposes and the action continues until all the aluminium has been transformed into alumina, and this even after the mercury has been removed. Somewhat similar results are obtained with magnesium and mercury.

These experiments with activated aluminium or activated magnesium are very curious, and are no doubt examples of electrolysis indicated in connection with the rusting of iron (p. 439); the effect of the copper zinc couple (p. 452); and amalgamated zinc (p. 408). If the aluminium is below 98 per cent, purity, it is said that the amalgam will not attack water appreciably at ordinary temperatures. In virtue of its intense electropositive character, aluminium is a powerful reducing agent. Thus, carbon di- and mon-oxides are reduced to carbon; $3\text{CO}_2 + 4\text{Al} = 2\text{Al}_2\text{O}_3 + 3\text{C}$; and $3\text{CO} + 2\text{Al} = \text{Al}_2\text{O}_3 + 3\text{C}$. Aluminium powder, when intimately mixed with many metallic sulphides or oxides—manganese, chromium, tungsten, uranium, iron, etc.—along with some flux, say, fluorspar, when ignited, reduces the oxides or sulphides to the metal. *E.g.* with pyrites, it forms aluminium sulphide and metallic iron. The mixture becomes very hot during the reaction and a temperature of $3000^\circ\text{--}3500^\circ$ is sometimes attained. The heat of this reaction can be utilized for softening and welding iron rails, steel castings, etc., when an intense local heat is needed. The rails to be welded are packed in a mixture of iron oxide and aluminium powder together with a special cement to make the mass compact. When the mass is ignited, it burns and heats the rails to a temperature high enough to weld the metals together. The mixture of aluminium powder with various metallic oxides is sold as "thermite," and the process is called H. Goldschmidt's or the aluminothermic process. The high temperature, and the extreme portability of the thermite process of welding and reducing metal oxides, is well emphasized by the remark of W. Ostwald: "Dr. Hans Goldschmidt has given the industrial world a blast furnace and smithy for the waistcoat pocket."

Aluminium, when heated with the halogens and with nitrogen, forms halides and nitride respectively. Aluminium dissolves slowly in cold dilute hydrochloric acid, and rapidly in hot, the concentrated acid giving an aqueous solution of aluminium chloride, and hydrogen gas. Nitric acid, dilute or concentrated, acts so slowly that nitric acid is usually said to have "no action on aluminium." Sulphuric acid has very little action in the cold, but the hot concentrated acid converts it into aluminium sulphate with the evolution of sulphur dioxide. Aluminium is rapidly dissolved by sodium and potassium hydroxides with the evolution of hydrogen and the formation of potassium aluminate: $2\text{Al} + 6\text{KOH} = 2\text{Al(OK)}_3 + 3\text{H}_2$. Organic acids (*e.g.* acetic acid) are almost without action on the metal at ordinary temperatures, but they are said to have an appreciable solvent action in the presence of sodium chloride. Salt solutions, *e.g.* sea-water, rapidly corrode the metal.

Atomic weight.—The atomic weight of aluminium has been determined by roasting the metal to convert it into the oxide—0.4168 gram of metal furnished 0.7850 gram of the oxide; this shows that if oxygen = 16 be the standard, the combining weight of aluminium is 26.99; and by the method of Fig. 34, 0.7314 gram of aluminium, after treatment with sodium hydroxide, furnished 907.9 c.c., that is, 0.0816 gram of hydrogen; consequently, if hydrogen be unity, aluminium has a combining weight of 26.89. The combining weight has also been determined by analyses of aluminium chloride, aluminium bromide, aluminium sulphate, and ammonia alum. The results show that if oxygen be 16, aluminium has a combining weight between 26.99 and 27.43; the best representative value is probably 27.1.

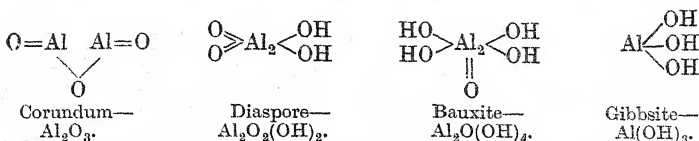
The approximate atomic weight by Dulong and Petit's rule, specific heat of aluminium: 0.2143, agrees with the value 27.1. The compounds of aluminium have a great inclination to form double molecules— AlCl_3 becomes Al_2Cl_6 , for example. Accordingly, aluminium was once considered to be quadri- or sexi-valent, and the two aluminium atoms were supposed to be in direct contact with one another $\text{Cl}_3\equiv\text{Al}-\text{Al}\equiv\text{Cl}_3$ or $\text{Cl}_3\equiv\text{Al}\equiv\text{Al}\equiv\text{Cl}_3$. The vapour densities of the aluminium halides indicate that 27.1 is the atomic weight, and that the metal is trivalent. The trivalency of aluminium is also in agreement with the isomorphic law applied to compounds of gallium, indium, thallium, titanium, vanadium, chromium, manganese, etc.

Alloys.—See aluminium bronze discussed in connection with the alloys of copper. *Magnalium* is the trade name of an alloy containing about 90 per cent. of aluminium with 2 to 10 per cent. of magnesium and small quantities of other metals. It has a specific gravity ranging from 2.40 to 2.57, and melts between 1184° and 1248° . It is claimed that it has all the good properties of aluminium, without its disadvantages. For instance, it is claimed that the tensile strength of magnalium is greater than aluminium, and the alloy can be turned in a lathe. Magnalium, under some conditions, is also less corroded than aluminium, zinc, copper, or brass. Hence it is replacing aluminium for many purposes. *Aluminium amalgam* is made by adding aluminium filings to a $\frac{1}{2}$ per cent. solution of mercuric chloride for a couple of minutes, and washing the product with alcohol and ether. As indicated above, the amalgam decomposes water at the ordinary temperature, liberating hydrogen and forming aluminium hydroxide. The mercury seems to retard the formation of a thin coating of oxide on the surface of the metal which would prevent further oxidation. Aluminium amalgam is a valuable neutral reducing agent because neither acid nor alkali is needed for the reaction.

Uses.—Large quantities of aluminium powder mixed with oil are used as paint for steam-pipes, and other metal objects exposed to heat or the weather. Aluminium is used for cooking utensils; the metallic parts of military outfits; certain parts of airships, etc.; precision instruments; surgical instruments; and as an ornamental metal for interior decoration; and artistic objects, trinkets, etc. The wire is used as a conductor of electricity, because aluminium wire, though thicker than copper for a given conductivity, is not so heavy and does not strain the supports so much. Aluminium is used as a reducing agent in the production of certain metals—chromium, etc.—and in the manufacture of "thermite." The formation of oxides during the melting of many metals is prevented if a little aluminium be present, hence aluminium—0.16 to 0.05 per cent.—is commonly added to molten steel as it comes from the Siemens-Martin's, or Bessemer's furnace. This enables castings to be made more free from "blow holes." Aluminium is difficult to solder, and therefore the parts of large articles are commonly welded together by autogenous soldering.

§ 6. Aluminium Oxide and Hydroxide.

Three hydroxides of aluminium occur in nature, and they can be represented graphically, though empirically, by assuming that aluminium is a triad, and that alumina, or corundum, is approximately Al_2O_3 , or,



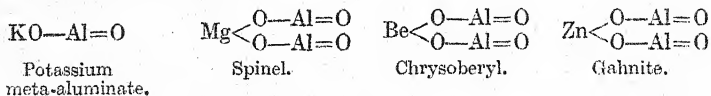
The gelatinous precipitate which is obtained when ammonia or an alkaline carbonate is added to a solution of an aluminium salt, is a colloidal aluminium hydroxide— $\text{Al}(\text{OH})_3$ —and it contains absorbed water, $\text{AlCl}_3 + 3\text{NH}_4\text{OH} = \text{Al}(\text{OH})_3 + 3\text{NH}_4\text{Cl}$. The curve obtained by measuring the rate at which the gelatinous hydrates lose water on a rising temperature does not show, in a convincing manner, the transition from one hydrate to the other. The vapour pressures of hydrates of this kind do not vary in steps corresponding with the existence of definite hydrates—p. 531—for the molecular ratio of the water can change continuously without abrupt changes. This seems to contradict the law of multiple proportions, and to get over the difficulty it is said that the water is not chemically combined, and substances of this kind are called *absorption compounds*. Aluminium hydroxide can be obtained in the sol (hydrosol) condition by dialyzing an aqueous solution of aluminium chloride, or a solution of freshly precipitated aluminium hydroxide in aluminium chloride or aluminium acetate; and in the gel (hydrogel) condition by adding some coagulating salt as indicated on p. 463. Freshly precipitated aluminium hydroxide dissolves easily in acids and in alkaline hydroxides. If the precipitate has stood a long time under water, or if the precipitate be dried, it dissolves very slowly in these reagents.

When aluminium hydroxide is precipitated in a solution containing a colouring matter, the latter is simultaneously precipitated, and the aluminium hydroxide with the absorbed colouring matter is called a lake. Advantage is taken of this property in dyeing cloth. The aluminium hydroxide is first precipitated in the fibres of the cloth, and the fabric is then immersed in the dye, and some of the dye is fixed by the aluminium hydroxide in the fibres. Hence, aluminium hydroxide is a dye-fixing agent or mordant. Dyes which stain the fibres directly need no mordant.

Aluminium hydroxide as a basic oxide.—Aluminium hydroxide is amphoteric, for it exhibits both feebly acidic and feebly basic properties. The existence of salts like aluminium chloride, aluminium sulphate, etc., is evidence of the basic qualities; and the feeble basic qualities of the hydroxide is indicated by the fact that salts with the weak acids—carbonic, hydrosulphuric, and sulphurous acids—do not exist in aqueous solution.

Aluminium hydroxide as an acidic oxide.—The feeble acidic properties of aluminium hydroxide are evidenced by the solubility of the hydroxide in alkalis and by the formation of salts—aluminates—with the strong bases. The hydroxide H_3AlO_3 is called orthoaluminic acid when it is desired to emphasize its acidic nature. The corresponding salts, $\text{Al}(\text{OM})_3$, $\text{Al}(\text{OH})(\text{OM})_2$, $\text{Al}(\text{OH})_2(\text{OM})$, are orthoaluminates: $\text{Al}(\text{OH})_3 + 3\text{NaOH} = \text{Al}(\text{ONa})_3 + 3\text{H}_2\text{O}$; $\text{Al}(\text{OH})_3 + \text{NaOH} = \text{AlO}(\text{ONa}) + 2\text{H}_2\text{O}$. These compounds remain behind when the solutions are evaporated to dryness. They also separate when alcohol is added to the solutions. Solutions of

the aluminates furnish precipitates of the respective aluminates when salts of the alkaline earths are added: $2\text{NaAlO}_2 + \text{CaCl}_2 = \text{Ca}(\text{AlO}_2)_2 + 2\text{NaCl}$. Salts corresponding with NaAlO_2 are considered to be derived from the monobasic acid HAlO_2 , called meta-aluminic acid, which corresponds with the ortho-acid less one molecule of water. The minerals *spinel*— $\text{MgO} \cdot \text{Al}_2\text{O}_3$, or magnesium meta-aluminate, $\text{Mg}(\text{AlO}_2)_2$; *chrysoberyl*— $\text{BeO} \cdot \text{Al}_2\text{O}_3$, or beryllium meta-aluminate, $\text{Be}(\text{AlO}_2)_2$; *gahnite*— $\text{ZnO} \cdot \text{Al}_2\text{O}_3$, or zinc meta-aluminate, $\text{Zn}(\text{AlO}_2)_2$; *hercynite*, $\text{FeO} \cdot \text{Al}_2\text{O}_3$, or ferrous meta-aluminate—are supposed to be meta-aluminates which can be represented by the graphic formulæ:



Pleonaste or *ceylonite*, $\text{MgO} \cdot \text{Fe}_2\text{O}_3$; *franklinite*, impure $\text{ZnO} \cdot \text{Fe}_2\text{O}_3$; *magnetioferrite*, $\text{MgO} \cdot \text{Fe}_2\text{O}_3$; *magnetic oxide of iron*, $\text{FeO} \cdot \text{Fe}_2\text{O}_3$; and *chromite*, $\text{FeO} \cdot \text{Cr}_2\text{O}_3$, are sometimes included in this group of minerals with the general name spinels, and general formula $\text{R}''\text{O} \cdot \text{R}'''\text{O}_3$, where R'' represents the dyad elements Fe, Mn, Mg, Be, Zn; and R''' the triad elements Fe, Mn, Al, Cr.

Aluminates.—The attempt to show that aluminium hydroxide has no acidic properties, and that the so-called aluminates are colloidal solutions or solid solutions, has failed. The electrical conductivity of solutions of sodium aluminate is characteristic of that for a monobasic acid. The aluminates are not very stable. Their aqueous solutions are strongly alkaline on account of hydrolysis. They are decomposed by carbon dioxide with the precipitation of aluminium hydroxide: $2\text{NaAlO}_2 + \text{CO}_2 + 3\text{H}_2\text{O} \rightleftharpoons 2\text{Al}(\text{OH})_3 + \text{Na}_2\text{CO}_3$. There is some doubt about the existence of an aluminium carbonate, although it is probable that a basic carbonate is produced which is probably formed by the hydrolysis of the normal carbonate. Ammonium chloride produces a precipitate of aluminium hydroxide owing probably to the immediate hydrolysis of an unstable ammonium aluminate: $\text{Al}(\text{ONa})_3 + 3\text{NH}_4\text{Cl} \rightleftharpoons \text{Al}(\text{ONH}_4)_3 + 3\text{NaCl}$; which is completely hydrolyzed by water: $\text{Al}(\text{ONH}_4)_3 + 3\text{H}_2\text{O} \rightleftharpoons 3\text{NH}_4\text{OH} + \text{Al}(\text{OH})_3$. The aluminium hydroxide so obtained is a pulverulent powder not gelatinous like that precipitated from the acid solution of aluminium salts; it is also much less readily dissolved by acetic acid. The alumina of commerce has usually been precipitated from sodium aluminate, and it generally contains some sodium carbonate due to the imperfect washing of the precipitate. An excess of hydrogen peroxide when added to a concentrated solution of potassium aluminate gives a white precipitate, which, when filtered and dried, liberates iodine from potassium iodide, and gives perchromic acid with an acidified solution of chromic acid. It seems to be aluminium peroxide. There are many analogies between the compounds of aluminium and beryllium—e.g. the formation of beryllonates and aluminates, etc.

It is interesting to note that "intermediate oxides"—that is, oxides which can act both as acids and bases—must necessarily have both properties *feebly* developed because, in the language of the ionic hypothesis, the acidic and basic qualities depend on the presence of H^+ and OH^- ions,

and both ions cannot be present in very great concentration in the same solution, owing to their tendency to unite and form water.

Aluminium oxide, Al_2O_3 .—Aluminium oxide, or alumina, occurs in nature as colourless crystalline corundum; and tinted with various metallic oxides as ruby, sapphire, amethyst, emery, etc. Alumina is prepared as a white powder by the ignition of aluminium hydroxide, aluminium nitrate, or ammonia alum. Alumina fuses at 1900° to 2010° , and it begins to volatilize appreciably at 1750° . Alumina is prepared in a crystalline condition by strongly heating a mixture of aluminium fluoride and boric oxide: $2\text{AlF}_3 + \text{B}_2\text{O}_3 = \text{Al}_2\text{O}_3 + 2\text{BF}_3$. Artificial rubies have been made by heating alumina with chromic oxide and boric oxide. The latter acts as a flux and gradually volatilizes, leaving behind the crystalline artificial "gem." When alumina is heated above about 800° an exothermal change takes place, the alumina changes in some way, for it then becomes almost insoluble in acids; its specific gravity rises rapidly from 2.8 to 4.0; and other physical properties change at the same time. The change is supposed to be due to the formation of an allotropic modification of alumina. Similar remarks apply to the effects of high temperatures on ferric, chromic, and many other oxides. When heated with reducing agents—potassium, sodium, calcium carbide, etc.—alumina is reduced to the metal. Bauxite is used in the manufacture of the so-called "bauxite bricks," and for lining the bed of basic open hearth furnaces. Good grade bauxite bricks soften between 1600° and 1800° . Fused bauxite or fused alumina is manufactured, sold under such trade-names as "alundum," "diamantine," "aloxite," "electrite," "borocarbene," etc., and used in the manufacture of abrasive and refractory materials. Fused alumina is also used as a high-grade refractory for muffles, etc.

§ 7. Gallium, Indium, and Thallium.

The rare metal gallium, Ga, was discovered by Lecoq de Boisbaudran in 1875 while studying a zinc blende from the Pyrenees, and named after Gallia, the Latin name of his country. Indium, In, is another rare metal discovered in 1863 by T. Reich and F. Richter in a zinc ore from Freiberg. Both elements were discovered by the spectroscope. The spark spectrum of gallium contains two violet lines, and indium has a characteristic bright indigo-blue line. The latter element was named from its prominent indigo spectral line. Gallium and indium metals are readily attacked by water, although indium slowly decomposes water at ordinary room temperatures. Gallium and indium are attacked by nitric acid, whereas aluminium under the same conditions appears to be passive. Gallium and indium are related to aluminium much as zinc is related to magnesium. Aluminium, gallium, and indium form oxides R_2O_3 . The hydroxides $\text{R}(\text{OH})_3$ have weak acidic and basic properties. All three elements form well-defined isomorphous ammonia alums. Unlike gallium chloride, GaCl_3 , and aluminium chloride, AlCl_3 , indium chloride, InCl_3 , can be obtained by the evaporation of the aqueous solution at 100° with relatively little hydrolysis. All three elements are trivalent, but indium also forms three chlorides: InCl , InCl_2 , and InCl_3 ; and gallium forms GaCl_2 and GaCl_3 .

Thallium, Tl, was discovered by W. Crookes in 1861 while studying the flue dust from a sulphuric acid chamber at Tilkerorde (Hartz

mountains). Thallium occurs associated with pyrites in zinc ores, and in the mineral *crookesite*, a copper selenide containing 16 to 18 per cent. of thallium, and 3 to 5 per cent. of silver. Thallium was discovered by the spectroscope. Its spectrum has a characteristic green line—hence its name is derived from the Greek *θαλλός* (*thallos*), a green twig. The physical properties of the metal closely resemble lead, and its compounds are related to aluminium, gallium, and indium compounds, much as gold and silver are related to the alkali metals; and mercury to zinc and cadmium. Thallium forms a series of thallous salts similar in properties to mercurous and silver salts. The thallous halides, for instance, like those of silver, are but slightly soluble in water. Mercurous, thallous and silver chlorides are soluble in a solution of sodium thiosulphate. Thallous sulphide, like the silver and lead sulphide, is but sparingly soluble in water; and thallous hydroxide and carbonate are fairly soluble in water. The thallic salts are not very stable, and they behave somewhat like the auric salts, so that while thallous oxide, Tl_2O , is a strong base, the trioxide, Tl_2O_3 , is a feeble base.

§ 8. The Relationships of the Boron-Aluminium Family.

The variations in the physical properties of these five elements with increasing atomic weight is indicated in the following table:

TABLE XLIX.—PROPERTIES OF THE BORON-ALUMINIUM FAMILY.

—	Boron.	Aluminium.	Gallium.	Indium.	Thallium.
Atomic weight . .	11.0	27.1	69.9	114.8	204.0
Specific gravity . .	2.45	2.7	5.9	7.4	11.8
Atomic volume . .	4.5	10.0	11.8	15.5	17.3
Melting point . .	2200°	657°	30°	176°	285°

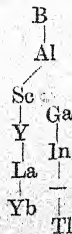
With the exception of boron and aluminium, the elements of this series are scarce and rare. The oxides of the metals from aluminium to thallium are obtained by adding ammonia to solutions of salts of the metals. The oxides dissolve in acids forming salts. The haloid salts are volatile. At about 100° thallic chloride breaks down into thallous chloride and chlorine. The chlorides readily form complex salts with other chlorides. The non-metallic characters predominate in boron, and it can almost be said that the properties of boron summarize those of the group, although any particularly striking family trait is wanting. Boric oxide exhibits strongly acidic and very feeble basic properties, for it combines with a few anhydrides—e.g. P_2O_5 and SO_3 —of the strong acids to form compounds in which boron acts as a base. Thus *boron phosphate*, BPO_4 , is a stable compound, so also is *boron sulphate*, $B_2(SO_4)_3$ (?), but is decomposed by heat or water. The other oxides are basic towards most acids, and the basic properties of the elements increase, and the acidic properties decrease with increasing atomic weights in passing from aluminium to thallium. Thus, aluminium is acidic towards strong alkalies, gallium and indium hydroxides are soluble in concentrated potash solutions, and there is

evidence of the formation of an unstable potassium thallate in the violet liquid which is produced when thallic hydroxide is suspended in a concentrated solution of potassium hydroxide, and treated with chlorine. The potassium hydroxide solution of indium oxide gives a precipitate of the oxide when boiled and the solution is decomposed by carbon dioxide. The trichloride, TiCl_3 , for instance, is partially hydrolysed by water. When thallium is trivalent, its compounds resemble the other members of the group; the thallic salts are readily reduced to thalious salts. The . . . ic salts are usually soluble in water and, excepting boron, the elements can replace aluminium to form alums—*e.g.* *gallium alums* $\text{M}_2\text{SO}_4 \cdot \text{Ga}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$; *indium alums*, $\text{M}_2\text{SO}_4 \cdot \text{In}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$. The grouping of thallium with aluminium is justified by the fact that thallium sesquioxide can replace alumina in the alums. Aluminium, indium, and thallium form complex sulphates with the alkali sulphates crystallizing with eight molecules of water; but when univalent, thallium behaves like silver and the alkali metals. The kinship of the thalious and alkali salts is illustrated by the fact that (i) thalious salts are stable bodies and usually isomorphous with the corresponding potassium salts; (ii) thalious sulphate can take the place of the alkali metal to form alums—*e.g.* $\text{Ti}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$; (iii) thalious monoxide, Ti_2O , dissolves in water, forming a corrosive solution which resembles that obtained with caustic alkali—it forms salts when treated with acids; absorbs carbon dioxide, forming thalious carbonate; colours red litmus blue; the carbonate and phosphate are soluble; the chloroplatinate, like that of potassium, is sparingly soluble; etc.; (iv) it forms a tri-iodide, isomorphous with potassium tri-iodide; and (v) it can replace potassium in potassium chromate, forming TiKCrO_4 . Thallium itself resembles lead.

§ 9. Scandium, Yttrium, Lanthanum, and Ytterbium.

	Scandium, Sc.	Yttrium, Y.	Lanthanum, La. (Neo).	Ytterbium, Yb.
Atomic weight	44.1	89.0	139.0	172.0

The four rare elements, scandium, yttrium, lanthanum, and neo-ytterbium, are related to the aluminium family. They are all triad elements. Their oxides are all of the type R_2O_3 , and their halogen compounds: RCl_3 , etc. The hydroxides are all basic and insoluble in alkaline hydroxides. The basicity increases in passing from scandium to ytterbium. Scandium, for example, is a very weak base, while lanthanum forms the hydroxide, with the evolution of heat, by the direct action of water on the oxide. All the elements form stable carbonates; the halides are non-volatile, and are but slightly hydrolysed by water. They form double sulphates— $\text{La}_2(\text{SO}_4)_3 \cdot 3\text{K}_2\text{SO}_4$ —with the alkali sulphates, but these compounds are not alums. A. von Welsbach (1906) and G. Urbain (1907) found that what was previously considered to be ytterbium is really a mixture of two elements—ytterbium proper or neo-ytterbium, and a new element, called by Urban, lutecium—from *Lutèce*, an old name for Paris. Lutecium has an atomic weight of 174. The relations between these elements and the aluminium family are often emphasized by the scheme shown in the margin.



Questions.

1. What weights of sodium, zinc, and aluminium respectively would be required to yield 168 litres of hydrogen measured at N. T. P. ? ($\text{Na} = 23$; $\text{Zn} = 65.4$; $\text{Al} = 27$).—*Aberdeen Univ.*

2. Give the preparation of pure aluminium oxide from bauxite and work out a similar method for preparing potassium chromate from chrome iron ore.—*Amherst Coll., U.S.A.*

3. Write the formulæ of the chief "alums" known. Point out their characters as a class. State and explain the principle they illustrate.—*London Univ.*

4. Compare the oxides of chromium with those of iron, of sulphur, and of aluminium.—*Massachusetts Inst. Technology, U.S.A.*

5. Mallet determined the atomic weight of aluminium (1) by estimating the hydrogen evolved by the action of caustic soda on the metal, when 5.2632 grams of aluminium gave enough hydrogen to make 5.2562 grams of water; (2) by analysis of the bromide, when 8.6492 grams of the bromide required 10.4897 grams of metallic silver for precipitation. Required the atomic weight from (1) and (2). ($\text{Ag} = 107.66$; $\text{Br} = 79.65$; $\text{O} = 15.96$).—*Science and Art Dept.*

6. Point out without using symbols the most marked features of resemblance and of difference in the chemical behaviour of the metals iron and aluminium and of their compounds.—*Cambridge Senior Locals.*

7. At 180° potash alum loses 43.67 per cent. of water. Point out the significance of this fact in deciding between the formulæ $\text{AlK}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ and $\text{Al}_2\text{K}_2(\text{SO}_4)_4 \cdot 24\text{H}_2\text{O}$.—*London Univ.*

8. What is the composition of borax, and to what class of salts does it belong? Mention some examples of other salts of similar composition. What is the action of a solution of borax on litmus?—*London Univ.*

9. Into what classes can the following oxides be placed with reference to their behaviour towards acids and bases—calcium oxide, aluminium oxide, manganese dioxide, sulphur trioxide, nitrogen peroxide? State the reasons for your classification.—*London Univ.*

10. What is an alum? Give the formulæ (a) ammonia alum, (b) potash alum, (c) chrome alum, (d) iron alum. Describe the preparation of a specimen of any one. What is the action of heat on (a)?—*London Univ.*

11. What metals are capable of yielding alums, and what do you infer from these facts as to the mutual relations of these metals?—*Board of Educ.*

12. If the heats of formation of the following compounds of sodium and aluminium are nearly :

Al_2O_3	2AlCl_3	2AlBr_3	2AlI_3	Na_2O	2NaCl	2NaBr	2NaI
195	162	134	87	100	194	182	148 Cals.

show that the principle of maximum work explains the facts that with sodium, chlorine will displace bromine; bromine, iodine; and iodine oxygen; while with aluminium, oxygen will displace chlorine; chlorine, bromine; and bromine, iodine.

13. What happens when the following substances are warmed with concentrated nitric acid—antimony, aluminium, arsenious oxide, phosphorus, hydrogen chloride, and ammonia?—*London Univ.*

14. Explain what happens when the following substances are submitted to electrolysis: sodium sulphate, potassium, hydrogen sulphate, brine, bauxite. Briefly indicate any practical details incidental to the processes.—*Cape Univ.*

15. Explain how a high temperature is obtained in the "aluminothermic" processes, and why so much heat can be made effective as compared with many other processes of combustion.—*Worcester Polyt. Inst., U.S.A.*

16. Aluminium is often regarded as a triad metal. State the facts on which that conclusion is based, and cite any evidence which seems to conflict with that view of the valency of the element.—*Science and Art Dept.*

17. Of the following compounds dissolved in water, state which will alter the colour of litmus to red or blue, and which are neutral: NaCl , Na_2CO_3 , $\text{Na}_2\text{B}_4\text{O}_7$ (borax), Na_2SO_4 , NaHSO_4 , Na_2HPO_4 , CaCl_2 , AgNO_3 , CuSO_4 , $\text{Bi}(\text{NO}_3)_3$.—*Science and Art Dept.*

18. State the general characters of the magnesium-zinc-cadmium family of elements, and indicate their relation to the alkaline earths on the one hand, and to aluminium on the other.—*Science and Art Dept.*

19. How is potash-alum prepared on a large scale? Illustrate its constitution by a formula on the assumption that aluminium is tervalent. Define the term "alum" as used in a generic sense.—*London Univ.*

CHAPTER XXXIV

THE PLATINUM METALS

§ 1. Review of the Platinum Metals.

THE "platinum metals" include platinum, Pt; iridium, Ir; osmium, Os; palladium, Pd; rhodium, Rh; and ruthenium, Ru. They occur in a metallic condition in gravels and sands associated together as mixtures or compounds along with magnetite, gold, chromite, etc., principally in that part of the Perm district in the neighbourhood of Nizhni Tagilsk drained by the Tura river and particularly its tributaries, the Iss and the Veeya, and Mount Blgodat in the Ural mountains (Russia), and in smaller quantities in California, Sumatra, Abyssinia, New Granada, Brazil, Australia, etc. The world's production amounts to about 6.5 tons per annum; and of this, about 6 tons comes from the Urals, and a little less than half a ton from the Choco district in Columbia (S. America), and other localities. The deposits in Russia, Columbia, Abyssinia, and Borneo are alone considered to be worth working. Some is said to occur at Wenden (Westphalia). The yield of platinum from other deposits is too small to be of commercial significance. The world's production of iridium amounts to about 5000 ozs. The platiniferous sands and gravels are washed as in the case of alluvial gold. "Platinum concentrates" consist of more or less rounded grains which sometimes show signs of cubic crystallization, flattened scales, and small nuggets containing approximately the following percentage composition:

Platinum.	Iridium.	Rhodium.	Palladium.	Gold.	Copper.	Iron.	Osmiridium.	Sand.
76.4	4.3	0.3	1.4	0.4	4.1	11.7	0.5	1.4

The platinum content may vary from 50-86 per cent., nuggets weighing up to 20 lbs. have been found on rare occasions. When the native metal contains a preponderating proportion of iridium (say up to 75 per cent.), it is called *platiniridium*. This variety is rare; it is white and harder and denser than platinum. *Iridosmium*, or *osmiridium* is a native metallic alloy containing approximately:

Platinum.	Iridium.	Rhodium.	Osmium.	Ruthenium.
10.1	52.5	1.5	27.2	5.9

with traces of palladium, copper, and iron.

The metals are greyish-white and lustrous. They all melt at a high temperature. They are not acted on by air or oxygen at ordinary temperatures. Osmium alone burns when strongly heated in air, forming the

tetroxide OsO_4 ; the others are scarcely affected, chemically, at any temperature. Palladium readily dissolves in hot nitric acid and very sparingly in hydrochloric acid, but the other metals are scarcely affected by hot acids. Aqua regia attacks osmium, forming osmium tetroxide, OsO_4 ; and it dissolves platinum forming the tetrachloride, PtCl_4 ; ruthenium is slowly dissolved, while iridium and rhodium are not appreciably attacked, unless in a very finely divided condition. The native platinum alloy stubbornly resists attack by chemical reagents—even aqua regia. The metals are readily reduced from their compounds, which fact probably accounts for their occurrence free in nature. The metals fall naturally into two groups with nearly equal atomic weights. The *light platinum metals* include ruthenium, rhodium, and palladium; the *heavy platinum metals* include osmium, iridium, and platinum. Palladium is related with silver, and platinum with gold, as indicated in Table L. Unlike the other platinum metals, osmium vaporizes without melting.

Extraction of the metals.—The gold can be removed from the platinum concentrates by the amalgamation process. About 1860, the platinum, contaminated more or less with other metals, was extracted by a smelting process; to-day, a wet process is used. Details of the process are a "trade secret." In a general way it may be said that the concentrates are digested with dilute aqua regia under pressure. The insoluble residue contains sand (gangue) and osmiridium. Osmium and ruthenium form volatile oxygen compounds which can be easily removed from the solution by distillation. The solution of crude platinum in aqua regia is evaporated to dryness and heated to 125° ; the residue is then dissolved in water; acidified with hydrochloric acid; and treated with ammonium chloride, when a precipitate containing platinum and iridium compounds is obtained from which the metals are obtained by ignition. Aqua regia dissolves the residual platinum, not iridium. The platinum can be obtained by precipitation as before. The mother liquid is worked over for palladium and rhodium. The further separation of the platinum metals from one another is a difficult and laborious operation because the properties of the metals are so much alike; and because the behaviour of the salts of one element is modified by the presence of others. Thus, iridium does not dissolve in aqua regia, but if iridium be alloyed with platinum, some iridium passes into solution when the alloy is digested in aqua regia.

TABLE L.—PROPERTIES OF THE PLATINUM METALS.

	Light.				Heavy.			
	Ru	Rh	Pd	Ag	Os	Ir	Pt	Au
Atomic weight	101.7	103.0	106.0	107.88	191.0	193.0	194.8	195.7
Specific gravity	12.26	12.1	11.9	10.6	22.47	22.38	21.45	19.31
Atomic volume	8.3	8.5	8.9	10.1	8.5	8.6	9.1	11
Melting point	c.2000°	1650°	1549°	962°	2300°	2000°	1755°	1064°
Boiling point	2780°	2750°	2820°	1955°	2950°	2850°	2650°	2204°
Valency	3, 8	2, 3, 4	2, 4	1, 2, 3	2, 3, 4, 8	2, 3, 4	2, 4	1, 3

History.—There is supposed to be a reference to platinum in Pliny's "Natural History" under the name "aluta." The term "platina del Pinto" for a white metal resembling silver, has been for a long time in general use by the Spaniards in South America. "Platina" is the diminutive form of the Spanish *plata*, silver, and "Pinto" has reference to the river where it was discovered. At one time its export from South America was forbidden by the Spanish Government, who ordered it to be thrown into the sea to prevent its being used for adulterating gold. In 1788, the Spanish Government bought it for about 8s. per lb., presumably for adulterating gold. It is now worth from £13 to £20 per troy oz. It was brought to Europe in 1735 by C. Wood, and W. Brownrigg and R. Watson described its properties in 1750. The plantina del Pinto of Choca (Colombia) attracted the attention of Antonio de Ulloa in 1748; J. J. Scaliger of Leyden, 1758, also referred to an infusible metal which has been taken to have been platinum. The metal attracted much attention at the time. Before 1823, most of the platinum in commerce came from South America. Platinum was discovered in Ural in 1819, and in 1824 Russia began exporting platinum; since that time, most of the platinum of commerce has come from that source.

Osmium and iridium were discovered by S. Tennant, 1802 to 1803; rhodium and palladium by W. H. Wollaston, 1803 to 1804; and ruthenium by K. Claus in 1845. All these metals were found during the study of native platinum. "Osmium" is named from the Greek *ὄσμη* (osme), a smell; "iridium" is named from the Greek *ἶρις* (iris), a rainbow, from the varying tint of its salts; "rhodium" is named from the Greek *ῥόδον* (rhodon), a rose, from the rose-red colour of its salts; "palladium" is named after the planet Pallas, discovered the same year as palladium, 1802; "ruthenium" is named after *ruthen*, for Russia.

§ 2. The Chlorides and their Complex Acids.

Platinum tetrachloride, PtCl_4 .—Platinum dissolves in aqua regia. If the solution of hydrochloroplatinic acid, H_2PtCl_6 , be evaporated to dryness, and the residue gently heated, a solution of the residue in hot water deposits reddish-brown crystals of $\text{PtCl}_4 \cdot 5\text{H}_2\text{O}$ on cooling. The anhydrous chloride, PtCl_4 , can be made by drying the crystals over sulphuric acid and warming them in a current of chlorine; or by heating hydrochloroplatinic acid in a current of hydrogen chloride between 165° and 200° .

If a stream of chlorine be passed over hydrochloroplatinic acid between 364° and 374° , the tetrachloride is formed; between 430° and 440° what appears to be the trichloride, PtCl_3 ; about 582° , platinum dichloride appears, and it has been reported that at the same temperature some monochloride is formed, but this lacks confirmation. Iridium chlorides IrCl_4 , IrCl_3 , IrCl_2 , and IrCl have been reported to be formed under similar circumstances respectively at 100° , 763° , 773° , and 798° . Chlorides of all six platinum metals of the type PtCl_4 are known. Palladium tetrachloride is not known in a free state, but double chlorides with potassium, etc., are known.

Chloroplatinates.—If platinum chloride be crystallized from a solution

acidified with hydrochloric acid, or if an aqua regia solution of the metal be evaporated a number of times with an excess of hydrochloric acid to drive off the nitric acid, reddish-brown deliquescent crystals of the complex acid $\text{H}_2\text{PtCl}_6 \cdot 6\text{H}_2\text{O}$ are formed. This substance—the “platinum chloride” of commerce—is really hydrochloroplatinic acid. The acid is dibasic, and it forms a characteristic series of complex salts—the chloroplatinates. Potassium chloroplatinate, K_2PtCl_6 , for example, is a yellow crystalline precipitate made by adding the acid to a solution of potassium chloride. While the solubilities of the normal alkali chlorides in water *increase* in passing from lithium to caesium, the chloroplatinates *decrease* in solubility in passing from lithium to caesium. For instance, 100 c.c. of water at 10° dissolves, in grams :

Li_2PtCl_6	Na_2PtCl_6	K_2PtCl_6	Rb_2PtCl_6	Cs_2PtCl_6
Very soluble	1.15	0.90	0.15	0.05

The solubility of ammonium chloroplatinate, $(\text{NH}_4)_2\text{PtCl}_6$, is 0.6 at 10° , and it thus comes between potassium and rubidium chloroplatinates. The fact that the sodium salt is fairly soluble in 80 per cent. alcohol, while the potassium salt is almost insoluble, enables a mixture of the chloroplatinates of sodium and potassium to be separated. The ammonium salt behaves like the potassium salt. During the electrolysis of ordinary salts—silver nitrate, potassium chloride, etc.—the metal is deposited on the cathode; with the chloroplatinates, some of the platinum migrates as a PtCl_6^- -ion to the anode compartment. Again, silver nitrate precipitates Ag_2PtCl_6 , not AgCl , thus confirming the deduction that “ PtCl_6^- ” is a bivalent complex acid radicle or that the solution of the acid furnishes the ions 2H^+ and PtCl_6^- on electrolysis. The constitution of the chloroplatinates will be discussed later.

Platinum dichloride, PtCl_2 .—If hydrochloroplatinic acid be heated between 250° and 300° , it furnishes grey granular powder of platinum dichloride, PtCl_2 , insoluble in water; at higher temperatures, it decomposes into platinum and chlorine. Platinum dichloride is also formed when platinum is heated to about 582° in chlorine gas. All six metals of the platinum series form salts of the type PtCl_2 . Dihydrated palladium dichloride, $\text{PdCl}_2 \cdot 2\text{H}_2\text{O}$, is obtained by the simultaneous action of hydrochloric acid and chlorine, or a little nitric acid, upon palladium. The solution deposits brownish-red crystals of the dihydrate on evaporation under reduced pressure. The anhydrous salt is formed if the crystals are warmed; and also by the action of chlorine on the warm sulphide. The garnet-red crystals of the anhydrous salt decompose at a red-heat, forming what has been reported as monochloride, PdCl , but this is not certain. Palladium iodide, PdI_2 , is precipitated as a black insoluble powder when potassium iodide is added to solutions of palladium chloride. This reaction is used sometimes for the separation of iodine from the other halogens, since the other halogen salts of palladium are soluble. It is difficult to precipitate the palladium quantitatively, because the iodide is very soluble in the presence of an excess of potassium iodide. Carbon monoxide unites with platinous chloride forming carbonyl platinous chlorides, $\text{CO} \cdot \text{PtCl}_2$; $2\text{CO} \cdot \text{PtCl}_2$; $3\text{CO} \cdot 2\text{PtCl}_2$; carbon monoxide also reduces a solution of the chloride even in the cold.

Chloroplatinites.—When platinum dichloride is digested with hydro-

chloric acid, it furnishes a reddish-brown solution which is supposed to contain hydrochloroplatinous acid, H_2PtCl_4 . The acid has not been isolated, but the salts—chloroplatinites—are formed by treating the solution with, say, potassium chloride. Potassium chloroplatinite, K_2PtCl_4 , forms rose-red crystals. The same salt is formed by reducing potassium chloroplatinate with moist cuprous chloride, CuCl . Potassium chloroplatinite is used in platinum printing in photography. This chloride is reduced to metallic platinum by the action of light on ferrous oxalate. Palladium forms a similar series of chloropalladites.

Platinum-ammonia compounds.—The platinum metals behave in a peculiar manner with ammonia. Thus, when ammonia is added to a solution of platinum tetrachloride, PtCl_4 , in hydrochloric acid, a green precipitate is formed. If the mixture be boiled, a green insoluble compound, $\text{PtCl}_2 \cdot 4\text{NH}_3 + \text{H}_2\text{O}$, called Magnus' green salt, is formed and $\text{PtCl}_2 \cdot 2\text{NH}_3$ remains in solution. If the precipitate be heated to 250° , a yellow crystalline substance sparingly soluble in water is formed, $\text{PtCl}_2 \cdot 2\text{NH}_3$. Both compounds can be oxidized with chlorine to $\text{PtCl}_4 \cdot 2\text{NH}_3$. These two compounds may be taken to represent two well-defined series of platinum ammonia compounds. One series is derived from PtCl_2 and the other from PtCl_4 . These will be discussed very shortly.

A solution of platinum salts give a brown precipitate of the sulphide when treated with hydrogen sulphide, and the precipitate is soluble in alkali sulphide solution. Platinum salts also give a yellow crystalline precipitate with ammonium or potassium chloride. Most platinum salts decompose when ignited, and the metal is obtained on washing away the soluble products of decomposition. The metal itself is not affected by treatment with the mineral acids, although it is dissolved by aqua regia. Most platinum salts in solution are reduced by reducing agents.

§ 3. The Oxides and Hydroxides.

When a solution of potassium chloroplatinite or of platinumous chloride is treated with an alkaline hydroxide, platinumous hydroxide, $\text{Pt}(\text{OH})_2$, is precipitated as a black powder. It is soluble in the haloid acids—hydrochloric and hydrobromic acids—and in sulphurous acid, but not in the other oxy-acids, and thus forms the corresponding platinumous salts. The hydroxide is decomposed into the metal and platinum dioxide, PtO_2 , by boiling alkaline hydroxides: $2\text{Pt}(\text{OH})_2 = \text{PtO}_2 + \text{Pt} + 2\text{H}_2\text{O}$. When gently ignited, platinumous hydroxide forms the corresponding platinumous oxide, PtO , as a dark powder insoluble in water and in most acids. It is doubtful if IrO has been made. Palladium black warmed in air furnishes a dark bluish-green palladous oxide, PdO , which decomposes into the metal and oxide at a higher temperature.

When a boiling solution of potassium hydroxide is added to a solution of platinum tetrachloride, and the precipitated platonic hydroxide, $\text{Pt}(\text{OH})_4$, is washed with acetic acid to remove the potash, a yellowish powder is obtained which dissolves in acids, forming platonic salts, and in bases forming a series of salts called the platينات. For instance, with sodium hydroxide, yellow crystals of sodium platinate, $\text{Na}_2\text{O} \cdot 3\text{PtO}_2 \cdot 6\text{H}_2\text{O}$, are

obtained. Hence platinic hydroxide is an acidic and a basic compound. Platinic hydroxide, $\text{Pt}(\text{OH})_4$, is a type of similar compounds formed by the whole six of the platinum metals. The hydroxides when heated form dark grey powders of the dioxide—e.g. PtO_2 , IrO_2 , etc.

Ruthenium, osmium, iridium, and rhodium form sesquioxides: Ru_2O_3 ; Os_2O_3 ; Ir_2O_3 ; Rh_2O_3 . Ruthenium and osmium form compounds corresponding with the trioxides, RuO_3 and OsO_3 . A more or less impure IrO_3 has been made. Thus, potassium ruthenate, K_2RuO_4 ; and potassium perruthenate, KRuO_4 , call to mind potassium manganate and permanganate. Ruthenium and osmium also form tetroxides of the type RuO_4 and OsO_4 respectively. These compounds represent the highest known state of oxidation of any single metal. The nearest approach to this state of oxidation occurs with perchloric and permanganic acids. The tetroxides dissolve in water but the solutions are not acid: (1) they are neutral to litmus; (2) do not decompose carbonates; and (3) form crystalline salts. The acids show no signs of hydrolysis. The term "osmic acid" for osmium tetroxide is thus a misnomer. Both tetroxides, RuO_4 and OsO_4 , melt at a low temperature, about $+40^\circ$, and boil at about 100° giving irritating vapours. Osmium tetroxide vapours are very poisonous, and seriously injure the eyes. They decompose on further heating into the dioxide and oxygen. The solutions are reduced by organic matters and the finely divided metal is precipitated. The doubt whether, say, osmium is really octovalent in the tetroxide was lessened when osmium octofluoride, OsF_8 , was obtained with a vapour density of 355—theory for OsF_8 being 343—by the action of fluorine on osmium. *En passant*, it may be added that osmium tetrafluoride, OsF_4 , and osmium hexafluoride, OsF_6 , are formed at the same time. The tetrafluoride predominates when the action occurs at about 100° , and the other two fluorides, when the action occurs at about 250° . The octofluoride is separated in lemon-yellow crystals, melting at 34.5° , by cooling the product *in vacuo* in liquid air.

§ 4. The Properties and Uses of the Platinum Metals.

The progress which has been made and the considerable exactness actually attained in chemical analysis have been owing in a great measure to the discovery of platina. Without the resources placed at the ready disposal of chemists by this invaluable metal, it is difficult to conceive that the multitude of delicate analytical experiments which have been required to construct the fabric of existing knowledge could have ever been performed.
—J. F. W. HERSCHEL.

Platinum is a greyish white metal with a brilliant lustre. It is harder than copper, silver, or gold. It is ductile and malleable, and usually comes on the market in the form of foil or wire. Platinum has also the valuable quality that it softens like iron before melting, so that like iron it can be welded. Platinum melts between 1750° and 1755° ; and boils at about 2450° . Platinum and rhodium do not volatilize appreciably at 900° , but at 1300° volatilization can be detected. Appreciable quantities of the metals palladium, iridium, and ruthenium volatilize at 900° , and at 1300° the effect is very marked. Iridium is readily oxidized to a volatile sesquioxide when heated just below 1000° ; but it is

doubtful if platinum is oxidized below 1300° . Molten platinum, like molten silver, absorbs oxygen which is given off as the molten metal cools, hence it is liable to "spitting." Platinum is not attacked by pure hydrofluoric, hydrochloric, nitric, and sulphuric acids. It is readily dissolved by aqua regia and by solutions containing chlorine: see "Gold." When platinum is alloyed with silver, copper, lead, zinc, etc., it is attacked and partly dissolved by nitric acid, probably forming a platinum nitrate.

The high fusing temperature, and the fact that platinum is not attacked by air and strong acids enables it to be used in the manufacture of apparatus—dishes, crucibles, stills, etc.—for many chemical operations which could not be readily performed with apparatus made from other available metals. The unfortunate steady advance in the price will lead to the use of gold crucibles for many purposes. The analysis of many minerals could not be so readily conducted as at present if it were not for the valuable qualities of platinum.

Platinum is attacked by alkalis, nitrates, cyanides, and phosphates under reducing conditions. With phosphorus it forms platinum phosphide; with sulphur, platinum sulphide, PtS ; with sulphur and dry alkali, platinum disulphide, PtS_2 ; with arsenic, platinum arsenide, Pt_2As_3 . An arsenide called *sperrylite*, $PtAs_2$, associated with nickel sulphide occurs at Sudbury (Ontario). Platinum also alloys directly with metals like lead, silver, zinc, etc., but not mercury. Hence platinum crucibles must not be heated with these metals. Carbon alloys with platinum forming a brittle platinum carbide, and hence platinum crucibles must not be heated in a smoky flame.

Platinum has nearly the same coefficient of expansion as glass, and platinum wires can be fused in glass so as to make gas-tight joints. Platinum is also a good conductor of electricity, and large quantities are used in the electric light industry. Short pieces of platinum wire are fused into the glass at the base of the bulb, and connected with the filament inside. The bulb is then exhausted and sealed. The platinum wires outside are then put in communication with the wire carrying the current. The filament is thus heated under reduced pressure to form the "incandescent electric light." Platinum is used in dentistry, photography, in jewellery, and in making scientific and surgical instruments, etc. One-third of the world's supply of platinum is said to be used in dentistry, and another third in electro-technical work. The stock of platinum is gradually dissipated by wear in the laboratory, etc.; and, owing to the very limited occurrence of this metal, the reserves must also be approaching exhaustion. The demand for platinum outstrips the supply, and this naturally explains the enormous increase in price. One of the most serious problems in certain branches of technical chemistry, and in electro-technical work is based on the limited occurrence of this metal.

Platinum-iridium alloys are hard and elastic; malleable and ductile; and less fusible than platinum. If more than 20 per cent. of iridium be present, the alloys are exceedingly difficult to work. An alloy of 10 per cent. iridium and 90 per cent. platinum was chosen by the International Committee on Weights and Measures for preserving the standards of length and weight. Platinum-iridium wire with platinum wire are used as thermo-couples for temperatures up to 1000° ; and platinum-rhodium wires are used with platinum in a similar way for temperatures up to 1400° .

Commercial platinum has 2 per cent. of iridium, and it appears to gradually lose this constituent when heated to a high temperature. The result is that platinum crucibles made from commercial platinum lose in weight every time they are heated for some time in the gas blowpipe. This is a source of annoyance. The high fusing temperature of osmium has led to its use for the manufacture of filaments for incandescent electric lamps—"osmium lamps." These lamps give twice the illuminating power of the carbon filament lamp with an equal consumption of energy for an equal time. The glass bulb also remains unblackened even to the end of the life of the lamp; but other metal filament lamps have displaced the osmium lamp. An alloy of iridium and osmium is used for tipping gold nibs on account of its hardness. The alloy is called *iridosmine* or *osmiridium*. Palladium is used for absorbing hydrogen, for the detection of carbon monoxide, and for the separation of iodine as indicated above. Osmium tetroxide is used for straining and hardening organic tissues in histology.

When platinum is precipitated from solutions of the tetrachloride by reducing agents, a velvet black powder called platinum black is obtained; when ammonium chloroplatinate is calcined, the metal remains behind as a spongy mass called spongy platinum; and if asbestos be soaked in a solution of platinum chloride and ignited, the asbestos permeated with platinum is called platinized asbestos. Platinum sponge, platinum black, and platinized asbestos absorb large quantities of oxygen gas, and they can then be used as oxidizing agents. Platinum black can absorb 100 times its volume of oxygen and 110 times its volume of hydrogen. Palladium black absorbs about 900 times its volume of hydrogen. This property of occluding gases is shown in a less marked degree by iron, nickel, and cobalt, as well as by copper, gold, and silver. Spongy platinum will cause a mixture of hydrogen and oxygen to unite with explosion; spongy palladium without explosion. A jet of hydrogen directed on to finely divided platinum will cause the platinum to glow and finally ignite the jet of gas. Alcohol dropped on to iridium black takes fire. Similarly coal gas can be ignited by spongy platinum, and this property is utilized in making the so-called "self-lighting Bunsen burners." The catalytic properties of the finely divided platinum metals are used in some industries for promoting chemical changes, e.g. the contact process for sulphuric acid.

§ 5. Werner's Views on Valency, and on the Constitution of Molecular Compounds.

Sooner or later, theory, guided by observation, will disclose the most profound mysteries, and lift up the veil which prevents one seeing the disposition of atoms in the molecules.—M. A. GAUDIN.

The hypothesis that the valency of multivalent atoms is a pre-existing force acting with definite units of affinity is as unfounded as it is unnecessary.—A. CLAUS.

The attempt to distinguish molecular from atomic compounds, by structural formulæ based upon ordinary valencies deduced from the manifestations of the simple "atomic compounds," discussed in previous chapters, has not been successful. Some of the elements involved in the formation of the molecular compounds appear to manifest higher valencies

than the numbers deduced from the simpler, more numerous, and more stable compounds. A. Werner's hypothesis—1893 *et seq.*—seems to give a clearer insight into the constitution of double salts, complex salts, crystalline hydrates, etc., than any yet propounded. It is fully recognised that Werner's hypothesis has not assumed its final form, yet, as one of the fathers of philosophy has said: "We begin with doubts in order that we may end with certainties." Werner's hypothesis promises to banish the conception of molecular compounds as something specifically distinct from atomic compounds.

Residual Affinity.—C. L. Berthollet believed that: "there are pre-dominant affinities in substances which are the source of their characteristic properties; there are others which are inferior to these, but which nevertheless give rise to remarkable phenomena." According to Werner, when the combining capacity of an atom, as defined empirically by the theory of valency, is exhausted, the atoms still possess a "particular kind of affinity," which enables them to form molecular complexes, corresponding with Berzelius' "compounds of higher orders," p. 361. In other words, simple or primary molecules may possess a residual affinity which enables them to unite together and form more complex stable compounds. Thus, the sulphur atom in sulphur trioxide; the oxygen atom in water; the chlorine atom in hydrogen chloride; the nitrogen atom in ammonia; the gold atom in auric chloride; the platinum atom in platonic chloride, etc., all possess residual affinity which permits these molecules to unite additively with other molecules. Residual affinity appears to play a rôle similar to ordinary chemical affinity, but the new manifestations of valency differ from the manifestations of ordinary valency in that they bind entirely different radicles. For instance, residual affinity does not lead to the combination of univalent radicles as defined by the doctrine of valency. This does not mean that the mode of action of the two kinds of affinity is different, since both are in many ways similar. "Nevertheless," says Werner, "it appears at present desirable to preserve the difference because the doctrine of valency is yet in a transitional stage, and hence it is judicious to construct sharply defined concepts." Accordingly, Werner distinguishes two kinds of valency:

1. Chief or primary valency which represents those manifestations of chemical affinity which enable the combining capacities (valencies) of the elements to be expressed in terms of hydrogen atoms or their equivalents, *e.g.* $\text{Cl}-$, $\text{Na}-$, NO_2- , CH_3- , . . .

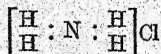
2. Auxiliary or secondary valency which represents those manifestations of (residual) chemical affinity which are able to bring about the stable union of molecules as if the molecules were themselves radicles able to exist as independent molecules, *e.g.* $\text{H}_2\text{O}-$, NH_3- , $\text{HCl}-$, CrCl_3- , . . .

In illustration, antimony pentachloride, SbCl_5 ; sulphur tetrachloride, SCl_4 ; phosphorus oxychloride, POCl_3 ; phosphorus pentachloride, PCl_5 ; and hydrogen cyanide, HCy , may all be regarded as compounds in which the primary valencies of the elements are exhausted, yet, in virtue of their residual affinity, compounds like $\text{SbCl}_5 \cdot 4\text{H}_2\text{O}$, $\text{SbCl}_5 \cdot 3\text{HCy}$, $\text{SbCl}_5 \cdot \text{SCl}_4$, $\text{SbCl}_5 \cdot \text{PCl}_5$, $\text{SbCl}_5 \cdot \text{POCl}_3$, etc., are readily formed.

Co-ordination number.—When the binding capacity of an elementary atom (primary valency) appears exhausted, Werner assumes that the atom can still link up with other molecules (secondary valency), and build

up more complex molecules. There is, however, a limit to the binding capacity of both primary and secondary valencies. The number which represents the joint effect of both the primary and the secondary valencies is called the "co-ordination number." The co-ordination number of an atom is the maximum number of atoms, radicles, or molecular groups— independent of their valencies—which can be directly linked with a central atom. The co-ordination number of an atom can be determined from compounds in which the maximum number of atoms, radicles, or molecular groups, linked with a central atom, are known. The co-ordination number of most atoms, curiously enough is six; in a few cases it is four; and with molybdenum and the addition products of the chlorides of the alkaline earths, the co-ordination number appears to be eight. The co-ordination number of some of the complex fluorides does not appear to be always six. The fact that the co-ordination number for so many elements is six, and is generally independent of the nature of the co-ordinated groups, has made Werner suggest that the number is decided by available *space* rather than affinity and that six is usually the maximum number which can be fitted about the central atom to form a stable system. Consequently the co-ordination number represents a property of the atom which enables the constitution of "molecular compounds" to be referred back to actual linkings between definite atoms. A molecular compound is primarily formed through the agency of secondary valencies; and, just as primary valencies determine the number of univalent atoms or their equivalent which can be linked to a central atom, so secondary valencies determine the number of molecules which can be attached to the central atom. The secondary valency is often active only towards definite molecular complexes, and hence the formation of additive compounds with other molecular complexes does not occur. Accordingly, the number of secondary valencies which are active towards different molecules is not always the same. To illustrate by example.

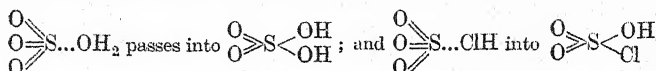
Ammonium chloride.—When it is desired to emphasize the distinction between primary and secondary valencies, Werner recommends using a continuous line for the former, and a dotted line for the latter. The nitrogen atom of ammonia, NH_3 , has an unsaturated secondary valency, and the hydrogen or chlorine atom in hydrogen chloride, HCl , has likewise an unsaturated secondary valency. The formation of ammonium chloride is therefore illustrated by the scheme: $\text{H}_3\text{N} + \text{HCl} = \text{H}_3\text{N} \cdots \text{HCl}$. The dotted line represents the auxiliary valency joining the hydrogen atom of HCl with the nitrogen atom of NH_3 . The co-ordination number of the nitrogen atom is here 4. It is not likely that one of the hydrogen atoms in ammonium chloride is "linked with a greater amount of affinity than the other three," and very probably, there is a state of equilibrium in which the affinity is distributed over all the hydrogen atoms, and a complex radicle is formed in which each of "the four atoms of hydrogen is united to the nitrogen atom by the same amount of affinity." Hence Werner writes the structural formula of ammonium chloride:



The practice of assuming an increased valency for nitrogen when ammonia unites with hydrogen chloride, says Werner, if consistently carried out,

would make antimony trivalent in antimony trichloride, and quinquivalent in $\text{SbCl}_3 \cdot \text{KCl}$, etc. Bivalent, Fe, in ferrous cyanide, FeCy_2 , would become decivalent in potassium ferrocyanide, K_4FeCy_6 .

Sulphuric acid.—The formation of sulphuric and chlorosulphuric acids by the union of sulphur trioxide with water and hydrogen chloride respectively is brought about by the secondary valencies as indicated in the schemes: $\text{O}_3\text{S} + \text{OH}_2 = \text{O}_3\text{S} \dots \text{OH}_2$; and $\text{O}_3\text{S} + \text{ClH} = \text{O}_3\text{S} \dots \text{ClH}$. *When one of the reacting molecules contains double-linked atoms, the secondary valencies may not be sufficiently strong to preserve the integrity of the new molecule, and the atoms of the addition product may be rearranged.* For example, this is the case with sulphur trioxide. Thus:

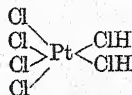
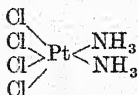


The change in the type of the compound is due to an intramolecular rearrangement, but such a change may not occur if the secondary valency is sufficiently strong and the addition product is stable.

Platinum-ammines.—What has just been suggested appears to be applicable with the analogous reactions: $\text{Cl}_4\text{Pt} + 2\text{NH}_3 = \text{Cl}_4\text{Pt} \dots (\text{NH}_3)_2$; and $\text{Cl}_4\text{Pt} + 2\text{KCl} = \text{Cl}_4\text{Pt} \dots (\text{ClK})_2$. Werner represents the two latter compounds by the formulæ:



All four chlorine atoms can be removed from the first compound without disturbing the ammonia molecules, and consequently the chlorine atoms do not act as intermediate links binding the NH_3 molecules to the platinum, as they would in, say, $\text{Cl}_2\text{Pt}(\text{Cl} \cdot \text{NH}_3)_2$. It is therefore inferred that the NH_3 molecules are directly attached to the platinum atom. Again, no difference has been detected in the chemical behaviour of the four chlorine atoms. This would not be the case if the two ammonia molecules were intermediate links between two of the chlorine atoms, and the central atom of platinum, as would be the case in, say, $\text{Cl}_2\text{Pt}(\text{NH}_3 \cdot \text{Cl})_2$. Hence, in all probability, the six groups are all attached directly to the platinum. Consequently, Werner writes:



The latter formula has been established in a similar manner to the former.

A study of the platinum-ammines shows that they can be arranged in two well-defined series. One series is derived from platonic chloride, PtCl_4 ; and the other from platinous chloride, PtCl_2 . The co-ordination number of the former is six; and of the latter, four. Thus:

1. *Platinum-ammines derived from platinous chloride— PtCl_2 .*

- | | |
|---|--|
| 1. Tetrammineplatinous chloride | $[\text{Pt}(\text{NH}_3)_4]\text{Cl}_2$ |
| 2. Chlorotriammineplatinous chloride | $[\text{Pt}(\text{NH}_3)_3\text{Cl}]\text{Cl}$ |
| 3. Dichlorodiammineplatinum (two isomers) | $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$ |
| 4. Potassium trichloroammineplatinitite | $[\text{Pt}(\text{NH}_3)\text{Cl}_3]\text{K}$ |
| 5. Potassium tetrachloroplatinite | $[\text{PtCl}_4]\text{K}_2$ |

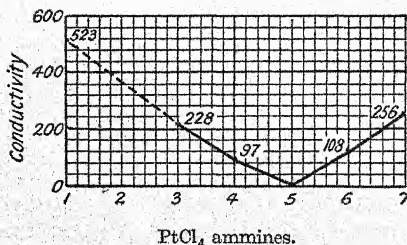
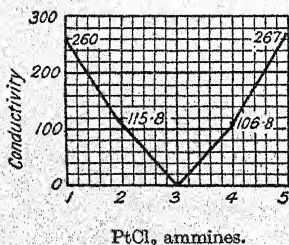
2. Platinum-ammines derived from platonic chloride— PtCl_4 .

1. Hexammineplatonic chloride $[\text{Pt}(\text{NH}_3)_6]\text{Cl}_4$.
2. Chloropentammineplatonic chloride $[\text{Pt}(\text{NH}_3)_5\text{Cl}]\text{Cl}_3$.
3. Dichlorotetrammineplatonic chloride $[\text{Pt}(\text{NH}_3)_4\text{Cl}_2]\text{Cl}_2$.
4. Trichlorotriammineplatonic chloride $[\text{Pt}(\text{NH}_3)_3\text{Cl}_3]\text{Cl}$.
5. Tetrachlorodiammineplatinum (two isomers) $[\text{Pt}(\text{NH}_3)_2\text{Cl}_4]$.
6. Potassium pentachloroammineplatinate $[\text{Pt}(\text{NH}_3)\text{Cl}_5]\text{K}$.
7. Potassium hexachloroplatinate $[\text{PtCl}_6]\text{K}_2$.

The simple or compound radicles which form the complex represented within the square brackets, are directly united with the central atom of platinum. The complex takes part in chemical reactions as if it were one individual radicle. The basic ammonia in the complex can be successively replaced by acidic radicles— Cl , Br , NO_3 , CO_3 , SO_4 , etc.—until the complex becomes acidic instead of basic and the platonic ammine complexes finally pass from electro-negative (basic) radicles to electro-positive (acidic) radicles. Potassium chloroplatinate is the limit of the platonic amines and potassium chloroplatinite is the limit of the platinitous ammine. It will be obvious that an enormous number of derivatives are conceivable.

Nomenclature of the metal amines.—Werner's system of naming the metal-ammonia compounds has been almost universally adopted. The constituents of the complex are taken first; and of these, the acid radicles with the suffix "o" come first; then follow any groups which behave like ammonia, e.g. H_2O is called *aquo*; NO_2 , *nitrito* or *nitro*; NO_3 , *nitrate*; CO_3 , *carbonato*; SO_3 , *sulphito*; SO_4 , *sulphato*; SCy , *thiocyanato*; etc. And lastly, preceding the metal itself, the ammonia molecules are designated "ammines," and spelt with a double "m" to distinguish the word from the "amines" or substituted ammonias. The prefixes di, tri, . . . indicate the number of each. The whole is written as one word. Examples appear in the above list.

The valency of the metal amines.—The valency of the complex is numerically equal to the difference between the ordinary valency of the central atom and the number of negative (acidic) elements or groups attached to the metal. Thus, the normal valency of platinum in the



FIGS. 247 AND 248.—The Electrical Conductivities of the Platinum-ammine Derivatives. (The abscissae refer to the numbers of the above tables.)

first of the above series is 4, hence the valency of $[\text{Pt}(\text{NH}_3)_5\text{Cl}]$, with one negative (acidic) group " Cl ," will be 3; this means that the complex in question acts as a trivalent electropositive (basic) radicle; and it can unite with three univalent electro-negative (acidic) radicles. The valency of $[\text{Pt}(\text{NH}_3)\text{Cl}_5]$ with five negative (acidic) groups, " Cl ," will be -1.

This means that the complex under consideration will act as a univalent electro-negative (acidic) radicle, and it can accordingly unite with one electropositive (basic) radicle like potassium, sodium, etc. If the valency of the acidic radicles in the complex are numerically equal to the normal valency of the central atom, the complex will be nullvalent. This is the case, for instance, with the complex $[\text{Pt}(\text{NH}_3)_2\text{Cl}_4]$.

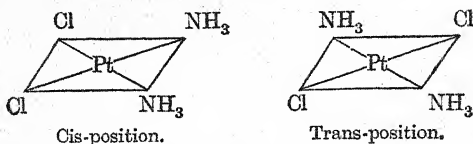
Properties of the metal ammines.—The nullvalent ammine bases are non-electrolytes, they do not conduct electricity. In the other ammine bases the complex forms one ion which is either electropositive (basic), and therefore a cation, or else electronegative (acidic), and therefore an anion. During electrolysis, the components of the complex are not disturbed. The molecular conductivities of the ammine derivatives of PtCl_2 at 1000 litres dilution are indicated in Fig. 247; the conductivities of the derivatives of PtCl_4 are indicated in Fig. 248; the second derivative, $[\text{Pt}(\text{NH}_3)_5\text{Cl}]\text{Cl}_3$, is not known. The electrical conductivities of these compounds correspond with the "ions" obtained on electrolysis, as indicated by Werner's formulæ in the above tables.

The hydroxides are strong bases, and in that respect resemble the caustic alkalis, for when boiled with ammonium salts, they expel ammonia; they absorb carbon dioxide from the air to form carbonates; they precipitate the hydroxides of aluminium, iron, etc., from solutions of their salts; and in the case of the aluminium salts, the precipitate is soluble in an excess of the base as is the case with caustic alkalis. Platinum sulphide is not precipitated from solutions of the ammines by hydrogen sulphide; the ammonia is not expelled by alkali hydroxide, nor by chlorine; the chlorine is but partially precipitated from the chlorides; etc. As a matter of fact, the analytical reactions of each base are characteristic of the complex as a whole, and of the radicles associated with the complex—indicated outside the square brackets in the above formulæ. Thus, with the cobalt ammines, cobalt trivalent, all the bromine in $[\text{Co}(\text{NH}_3)_6]\text{Br}_3$ is precipitated by silver nitrate; in $[\text{Co}(\text{NH}_3)_5\text{Br}]\text{Br}_2$ only two-thirds of the total bromine is precipitated by silver nitrate; and in $[\text{Co}(\text{NH}_3)_4\text{Br}_2]\text{Br}$ only one-third of the total bromine is precipitated; and in $[\text{Co}(\text{NH}_3)_3\text{Br}_3]$ none of the bromine is precipitated.

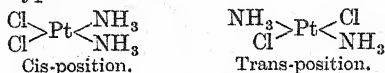
Isomerism of the ammines.—The phenomenon of isomerism occurs with some of the ammines. Thus, croceo-cobaltic chloride and flaveo-cobaltic chlorides have the same ultimate composition, $[\text{Co}(\text{NH}_3)_4(\text{NO}_2)_2]\text{Cl}$. One forms golden-yellow solutions—*flavis* (golden-yellow), and the other saffron-coloured solutions—*κρόκος* (crocos), saffron. Similarly, there are two isomers of dichlorodiammineplatinum and two isomers of tetrachlorodiammineplatinum. In the former, $[\text{Cl}_2\text{Pt}(\text{NH}_3)_2]$, the four radicles are attached to the central atom of platinum in pairs. If the four groups were attached in space, say at the angular points of a regular tetrahedron, isomerism could hardly be expected because the four groups could be interchanged without altering their relations one with another.¹ Hence it is inferred that the groups are arranged about the central atom of

¹ It is not legitimate to assume that the absence of isomerism in methylene chloride, CH_2Cl_2 , "proves" that the four hydrogen and chlorine atoms are distributed in space uniformly about a central carbon atom. The fact that no isomerism has hitherto been discovered may also "prove" that there is only one stable configuration for the group CH_2Cl_2 .

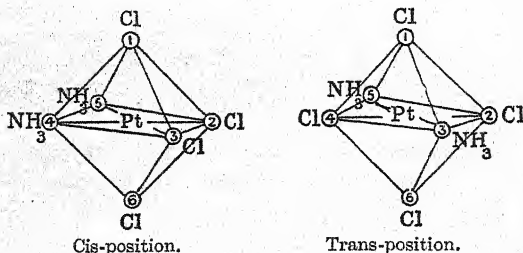
platinum in one plane. The resulting isomerism can be graphically illustrated by the schemes:



Or by the more usual type of formulæ :



The six associated units in co-ordination complexes appear to adopt the most symmetrical arrangement, and behave as if they were situated at the six vertices of a regular octahedron inscribed within the sphere of influence of the central atom. As Werner emphasized, the same forces are probably active in residual as in ordinary affinity, the difference being rather one of quantity than quality. The two isomers of $[\text{Pt}(\text{NH}_3)_2\text{Cl}_4]$ can be represented by :



The isomerism of other amines can be illustrated in a similar manner. If the corners of the octahedral figure be always numbered 1 to 6 as shown in the diagram, the relative positions of the acidic or basic radicles can be indicated by numbers. Assuming that the diagrams correctly represent the relative positions of the radicles, the cis-compound is called the 2.3-tetrachlorodiammineplatinum, and the trans-compound the 2.4-tetrachlorodiammineplatinum.

Cobalt-ammines.—Tervalent cobalt forms an important series of ammines. The ammonia groups are not here attached in the “loose” way typified by “ammonia of crystallization”—for they are in some way actually bound up so as to produce a complete change in the character of the salts. Thus, while the haloid is cobaltic chloride CoCl_3 can be readily detected by the usual analytical tests for chlorides, this is not always the case with the cobaltammines. Thus, with chloropentammine cobaltic chloride $\text{CoCl}_3 \cdot 5\text{NH}_3$, one halogen remains associated with the cobalt and ammonia during double decomposition with the salt, while two of them behave like the chlorine atom of chlorides. Thus, (a) they are precipitated by silver nitrate; and (b) two atoms alone are removed by treatment with concentrated sulphuric acid. The so-called luteo-cobaltic chloride $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$ is a hexamminecobaltic chloride, and it is formed in reddish-

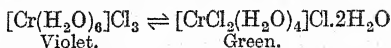
yellow prisms (*luteus* yellow) when a solution of cobaltous chloride containing ammonium chloride and ammonia is exposed to the oxidizing action of the air, or bromine, lead peroxide, etc. If an ammoniacal solution of cobaltous chloride be exposed to the oxidizing action of the air, roseo-cobaltic chloride, or aquopentamminecobaltic chloride, $[\text{Co}(\text{NH}_3)_5\text{H}_2\text{O}]\text{Cl}_3$, is formed; if this solution be warmed, purpureocobaltic chloride or chloropentamminecobaltic chloride, $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Cl}_2$, is formed; and when this is exposed to the action of nitrous acid, croceocobaltic chloride or nitritopentamminecobaltic chloride, $[\text{Co}(\text{NH}_3)_5\text{NO}_2]\text{Cl}_2$, is formed. Most of the 2000 known cobalt ammino-compounds can be referred to one of the following seven series. Let A represent NH_3 , H_2O , etc.; R represent the halogens, NO_2 , NO , CO_3 , Cy, etc.; and M a basic radicle univalent—K, Na, etc.

TABLE LI.—COBALTTAMMINES.

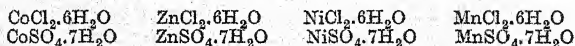
	Type.	Example—Werner's systematic name.	Formula.	Old name.
I.	$[\text{CoA}_6]\text{R}_3$	Hexamminecobaltic chloride	$[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$	Luteocobaltic chloride.
II.	$[\text{CoA}_5\text{R}]\text{R}_2$	Nitritopentamminecobaltic chloride	$[\text{Co}(\text{NH}_3)_5\text{NO}_2]\text{Cl}_2$	Xanthocobaltic chloride.
III.	$[\text{CoA}_4\text{R}_2]\text{R}$	Dichlorotetramminecobaltic nitrate	$[\text{Co}(\text{NH}_3)_4\text{Cl}_2]\text{NO}_3$	Chloropræseocobaltic nitrate.
IV.	$[\text{CoA}_3\text{R}_3]$	Trinitritotriamminecobalt	$[\text{Co}(\text{NH}_3)_3(\text{NO}_2)_3]$	—
V.	$[\text{CoA}_2\text{R}_4]\text{M}$	Ammonium tetranitritodiamminecobaltate	$[\text{Co}(\text{NH}_3)_2(\text{NO}_2)_4]\text{K}$	Erdmann's salt.
VI.	$[\text{CoAR}_5]\text{M}_2$	Potassium pentanitritodiamminecobaltate	$[\text{Co}(\text{NH}_3)(\text{NO}_2)_5]\text{K}_2$	—
VII.	$[\text{CoR}_6]\text{M}_3$	Potassium hexanitritocobaltate	$[\text{Co}(\text{NO}_2)_6]\text{K}_3$	Potassium cobaltinitrate.

Numerous derivatives of most of the classes are known, and an enormous number are possible, *e.g.* over a hundred members of the first class have been isolated. Potassium cobalticyanide, $[\text{CoCy}_6]\text{K}_3$, belongs to the seventh group. It is formed when a solution of a cobalt salt is warmed with an excess of potassium cyanide while exposed to the air. The cobaltous cyanide which first separates as a dirty brown precipitate forms the complex salt with the excess of potassium cyanide. The reaction is usually written $\text{CoCy}_2 + 4\text{KCy} = \text{K}_4\text{CoCy}_6$. According to Werner, this compound is $[\text{CoCy}_6]\text{K}_4$. This is oxidized to potassium cobalticyanide, $[\text{CoCy}_6]\text{K}_3$, on exposure to the air. Solutions of sodium hypobromite or hypochlorite have no action on the soluble salt. With nickel, a soluble salt, $[\text{NiCy}_4]\text{K}_2$, is formed under the same conditions, and this is decomposed by the sodium hypobromite, giving a precipitate of nickel hydroxide, $\text{Ni}(\text{OH})_2$. This is the principle of *Liebig's method of separating nickel and cobalt*. Cobalt salts when treated with acetic acid and potassium nitrite, and boiled for a few minutes, give a yellow precipitate of potassium cobaltinitrite, $[\text{Co}(\text{NO}_2)_6]\text{K}_3$; nickel salts do not give a precipitate under the same conditions because the potassium nickelinitrite is soluble. This is the principle of *Fischer's method of separating cobalt and nickel*.

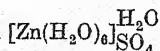
Chromium-ammines.—Similar remarks might be made, *mutatis mutandis*, about the ammines of iridium, chromium, etc. The two modifications of hydrated chromic chloride, $\text{CrCl}_3 \cdot 6\text{H}_2\text{O}$, are interesting—one is green, and the other violet. All the chlorine in the violet form, and one-third of the chlorine in the green form, can be precipitated by silver nitrate. Werner represents the violet form by $[\text{Cr}(\text{H}_2\text{O})_6]\text{Cl}_3$, and this, by losing two molecules of water, is transformed into the green hydrated chloride: $[\text{CrCl}_2(\text{H}_2\text{O})_4]\text{Cl} \cdot 2\text{H}_2\text{O}$. In the hydrated chloride, the chromium is present in the trivalent cation, and *all* the chlorine is in the anion; whereas in the green form only one Cl' ion is formed per molecule, the other two chlorine atoms are associated with the chromium to form a univalent radicle. The change is thus represented in symbols:



Hydrates.—We now come to the least satisfactory application of Werner's theory: crystalline magnesium chloride, $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$, is represented according to Werner's scheme by the formula $[\text{Mg}(\text{OH}_2)_6]\text{Cl}_2$, and similarly with the other salts crystallizing with six molecules of water. Werner thinks that the alums $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$, and some other hydrates with an abnormally large proportion of water, owe their existence to the addition of polymerized molecules of water; $(\text{H}_2\text{O})_2$, or H_4O_2 , is then attached as one molecule to the central aluminium atom, and alum becomes $[\text{Al}(\text{H}_4\text{O}_2)_6](\text{SO}_4)_2\text{K}$. When contrasting the chlorides and sulphates of the iron group, it seems curious that the former should crystallize with six and the latter with seven molecules of water. *E.g.*:



With these sulphates, *e.g.* $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$, Werner considers that one molecule of the water is attached to the sulphuric acid radicle: $\text{Zn}(\text{H}_2\text{O} \cdot \text{SO}_4) \cdot 6\text{H}_2\text{O}$, or $[\text{Zn}(\text{H}_2\text{O})_6]\text{H}_2\text{O} \cdot \text{SO}_4$,



in agreement with the great difficulty involved in driving off the last molecule of water, and with the fact that both potassium and ammonium sulphates are anhydrous, and yet when introduced into a sulphate with seven molecules of water, form a salt which crystallizes with six molecules—*e.g.* $\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$, *i.e.* $[\text{Fe}(\text{H}_2\text{O})_6]\text{SO}_4 \cdot (\text{NH}_4)_2\text{SO}_4$. Hence, in studying the hydrates, it is necessary to distinguish between acidic and basic water. While many of the hydrates have not yet been studied in the light of Werner's hypothesis, as Werner himself says, "These questions are for the future to decide; for the present, it is not possible to systematize all the hydrates."

Questions.

1. Whence and by what processes is platinum obtained? What are its physical properties and how is it affected by being alloyed with iridium? Point out briefly the utility for chemical investigation of platinum, glass, and india-rubber.—*New Zealand Univ.*

2. State precisely the methods you would suggest for the preparation of pure specimens of (a) sodium fluoride from calcium fluoride; (b) anhydrous magnesium

chloride from magnesium oxide; (c) a colloidal solution of silver from silver nitrate; and (d) potassium cobaltinitrite from cobalt.—*Calcutta Univ.*

3. Show by means of examples that valency depends not only upon the nature of the grouping atom but also on the nature of the atoms that are attached.—*Madras Univ.*

4. 0.2137 grm. of a substance gave 0.1371 grm. of CoSO_4 when evaporated to dryness with sulphuric acid; 0.1686 grm. gave 0.2026 grm. of AgCl when treated with silver nitrate; and 0.1542 grm. gave 58.08 c.c. of nitrogen at 19° and 721 mm. The compound was prepared by the action of sodium azide upon dichlorotetramminocobaltic chloride, $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]\text{Cl}$. What is the percentage composition of the substance under investigation and what is its probable formula?

5. How is platinum found in nature? Describe methods of obtaining it in a finely divided condition, and give one illustration of the use of the metal in an inorganic chemical preparation.—*Cape Univ.*

CHAPTER XXXV

THE OXIDES OF CARBON

§ 1. Carbon Dioxide—Preparation.

Molecular weight, $\text{CO}_2 = 44$. Melting point, -57° (under pressure); sublimes at -78.2° ; critical temperature, $+31^\circ$. Relative vapour density ($\text{H}_2 = 2$), 43.97; (air = 1), 1.529.

CARBON dioxide is produced when carbon burns in the presence of an excess of oxygen. Similar remarks apply to the combustion of many carbon compounds. If a beaker be held over a candle flame for a few minutes to collect some of the products of combustion, and lime-water be then poured into the beaker; or if a candle be allowed to burn in a cylinder loosely covered with a glass plate, and lime water be added, the formation of a turbid solution is strong circumstantial evidence that carbon dioxide is present. The production of a turbidity in clear lime-water is a characteristic test for carbon dioxide.

Laboratory methods.—Carbon dioxide is generally made in the laboratory by the action of hydrochloric acid upon calcium carbonate—marble, limestone, or chalk. The action is represented in symbols: $\text{CaCO}_3 + 2\text{HCl} = \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$. Calcium chloride is a by-product of the action. While hydrogen chloride expels carbon dioxide from the carbonates, if the gas be quite dry, this action does not occur. For example, Iceland spar, even at 130° , resists attack. Fragments of marble are placed in a Woulfe's bottle, Fig. 10, p. 58, with a quantity of water. Concentrated hydrochloric acid is added by means of a funnel tube. The gas can be collected, like hydrogen over water; or, unlike hydrogen, it can be collected by placing the delivery tube in an empty gas cylinder closed end downwards. The gas is so heavy that it displaces the lighter air upwards out of the jar—hence the term, *collecting the gas by the upward displacement of air*. The gas so collected may contain a little air. It is easy to test if the jar is full of gas because a lighted taper put down into the jar will be extinguished when it meets the carbon dioxide as completely as if it had been immersed in water.

If sulphuric acid be used in place of hydrochloric acid, the fragments of marble quickly become coated with a film of "insoluble" calcium sulphate which prevents further attack by the acid, and the evolution of gas practically ceases. If the powdered calcium carbonate be suspended in water, sulphuric acid may be used. Marble sometimes contains sulphur compounds which lead to the formation of sulphur dioxide, the carbon dioxide will then have a distinct smell of burning sulphur. By passing the gas through a wash-bottle containing a dilute sulphuric acid solution of

potassium permanganate, most of the sulphur dioxide will be removed from the gas. The gas, prepared as indicated above, may also contain steam and a little hydrogen chloride, as well as air. A fairly pure gas is made by the action of dilute sulphuric acid on lumps of "fused" sodium carbonate. In symbols: $\text{Na}_2\text{CO}_3 + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + \text{H}_2\text{O} + \text{CO}_2$. Sodium sulphate is a by-product of the reaction.

Large-scale manufacture.—Some carbon dioxide for manufacturing operations is made by the action of sulphuric acid on magnesite, whereby magnesium sulphate is obtained as a by-product. It is also made as a by-product when limestone or magnesite is burned in special kilns. In the latter case, the gas is purified by washing in towers, etc. For some purposes, too, carbon dioxide mixed with atmospheric nitrogen is made by passing air over red-hot coke. In the brewing industry, carbon dioxide is a by-product developed during the action of the yeast plant on sugar, or materials containing sugar. This gas is then washed and pumped into steel cylinders to be afterwards made into mineral waters.

Fermentation.—The formation of carbon dioxide during the process of fermentation is very curious. The fact can be illustrated by the following experiment. Shake 50 c.c. of treacle with 400 c.c. of water in a litre flask, *A*, Fig. 249, and add a few c.c. of yeast. Connect the flask with a glass delivery tube, *B*, in communication with an empty wash-bottle, *C*, and another wash-bottle, *D*, containing clear lime-water. The liquid in the flask, if kept warm, soon begins to bubble and froth—ferment. The whole apparatus is left on one side until the next lesson. The lime-water will then be quite turbid, and the contents of *D*

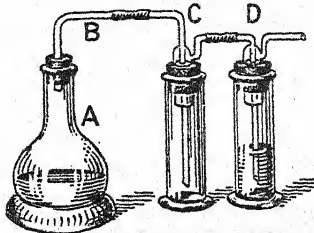


FIG. 249.—The Fermentation of Sugar.

may be tested for calcium carbonate as indicated below. The yeast plant during its growth decomposes the sugar— $\text{C}_6\text{H}_{12}\text{O}_6$ —solution forming carbon dioxide and ethyl alcohol— $\text{C}_2\text{H}_5\text{OH}$. The reaction, in symbols, is usually represented: $\text{C}_6\text{H}_{12}\text{O}_6 = 2\text{C}_2\text{H}_5\text{OH} + 2\text{CO}_2$. Besides ethyl alcohol, small quantities of other more complex alcohols are produced along with glycerol; succinic, acetic, formic, and other acids; aldehydes and esters are produced at the same time. Hence the equation just given represents only the main reaction, but there is quite an array of complex side reactions proceeding at the same time. The alcohol remaining in the flask *A* can be partially separated by fractional distillation. The last traces of water are removed from the distillate, boiling at about 80° , by distillation from freshly ignited lime, or recently fused potassium carbonate. Ethyl alcohol, when pure, boils at 78.4° (atmospheric pressure).

The "raising of bread" depends upon the expansion of bubbles of carbon dioxide by heat. The carbon dioxide is generated by the action of yeast on sugar or starch; or by the action of sodium bicarbonate ("baking soda") and a substance of acid reaction—tartaric acid, acid potassium tartrate (cream of tartar), etc. The formation of carbon dioxide during this reaction is shown by examining the gas developed when a mixture of cream of tartar and baking soda is moistened with water.

§ 2. The Properties of Carbon Dioxide.

The gas is invisible; when breathed through the nostrils, it gives a tingling sensation; and it has an acid taste.¹ If a jar of the gas be turned upside down for a minute, the heavy gas falls out and air takes its place. The gas is therefore heavier than air; and, as indicated above, it extinguishes the flame of burning bodies. These properties—the invisibility of the gas, its heaviness compared with air, and its property of quenching flame—enable many pleasing and attractive experiments to be arranged, and limited only by the ingenuity and skill of the experimenter.

Thus: (1) a long cardboard gutter leading into a glass trough in which a number of candles are burning can be arranged. The invisible gas, carbon dioxide, can then be poured from a large beaker down the gutter, when the candles will be extinguished just as if water had been poured down the gutter.

(2) A cylinder can be arranged with candles at different levels; the gas, when led into the cylinder, extinguishes the candles one by one as the gas rises in the jar.

(3) A wheel of stiff cardboard can be mounted upon two corks on a knitting needle journalled into a wooden standard; paper "buckets" can be attached to the periphery of the disc by joining the edges of the paper so as to form a hollow cone. When carbon dioxide is poured into the buckets, the wheel will rotate, if properly balanced, as if it were a "water wheel," moved by a stream of water.

(4) A cardboard box, or a light glass vessel can be counterpoised on a balance; the beam will be depressed when carbon dioxide is poured into the box, showing that the carbon dioxide is heavier than the air displaced.

(5) Soap bubbles blown with air can be floated on the carbon dioxide in a large dish.

(6) Petroleum burning in a shallow dish can be extinguished by pouring the gas over the dish; and a beaker of the gas poured over a lighted candle will quench the flame.

(7) Carbon dioxide can be ladled from a large cylinder by a small beaker tied on to a string. The action is analogous with the method of emptying a well of water by means of a bucket and rope. Finally, a candle will burn in the cylinder, showing that the carbon dioxide has been removed. The beakers of carbon dioxide can be emptied into another cylinder and the presence of carbon dioxide demonstrated in the usual way.

Carbon dioxide is an active agent in many portable fire extinguishers. Some contain sodium carbonate and sulphuric acid or alum solution so arranged that they can be mixed and the gases generated under pressure when desired. If enough carbon dioxide be forced on to the burning object the conflagration will be quenched. In "chemical fire engines" the pressure of the gas itself is utilized to force a stream of water on to the burning body.

Action on animals.—A mouse placed in a jar of the gas will be suffocated in a very short time. Carbon dioxide is not particularly poisonous. Its harmful effects are mainly due to suffocation (absence of oxygen). Small increases in the amount of carbon dioxide in air—say 2 or 3 per cent.—do no perceptible harm; 5–6 per cent. induces a rise of pulse and marked panting; 10 per cent., violent panting; above this the gas exerts a narcotic effect, and with 25 per cent. death will occur in a few hours; air with 50 per cent. carbon dioxide can be breathed a short time without fatal results. The ill-effects which attend breathing in a badly ventilated room are probably due to the presence of moisture and temperature. Moist still atmospheres hinder the proper evaporation of liquid from the

¹ The student should never smell or taste substances indiscriminately—smelling hydrogen cyanide might prove fatal; and the tasting of many substances would be equally disastrous.

surface of the body; the body temperature gets abnormally high, and the pulse greatly accelerated. An atmospheric temperature of 40° may not appear so distressing as one of 28° if the former permits ready evaporation of moisture, and the radiation of heat from the surface of the body, while the latter does not.

Solutions in water.—The solubility of carbon dioxide in water will be discussed later. Water dissolves about $1\frac{1}{2}$ times its volume of the gas at 0° and 760 mm. pressure; and about its own volume at 15° , 760 mm. In the manufacture of "soda water" the gas is dissolved by the water under great pressure—60–150 lbs. per square inch. The solution effervesces and froths when the pressure is withdrawn, owing to the escape of the carbon dioxide. Liquids bottled during fermentation—beer, champagne, kumiss, etc.—effervesce for the same reason. The same remarks apply to many mineral waters, e.g. the water of the Geyser Spring (Saratoga), at Niederselters (Hesse-Nassau), Seidlitz (Bavaria), etc., p. 178. Selters water is, in England, called "Seltzer water." The original "Apollinaris" is a natural water from a spring in the valley of Aar near the Rhine which has carbon dioxide in solution and an acid reaction, while the waters of Vichy Neuenahr, etc., have an alkaline reaction although the same gas is in solution. Apollinaris and other natural waters are imitated artificially and sold as "mineral waters"; the natural waters are often bottled and exported. Soda water, i.e. saturated with carbon dioxide, in the stomach acts as an aperient. Stale beverages are sometimes "revived" by saturating them with carbon dioxide under pressure.

Action of heat.—Carbon dioxide is fairly stable at high temperatures. When heated under atmospheric pressure, at 1300° , only 0.004 per cent. is decomposed; at 1400° , 0.14 per cent.; and at 1478° , 0.32 per cent. The action is a reversal of the oxidation process: $2\text{CO}_2 \rightleftharpoons 2\text{CO} + \text{O}_2$. So far as we can tell, two reactions are going on simultaneously: when the rate at which the carbon dioxide is decomposed is equal to the rate at which the carbon monoxide is oxidized, the two reactions are balanced, and the system is in equilibrium. Raising the temperature above 1300° augments the velocity of the process of decomposition.

Action of metals.—potassium, calcium, etc.

—If the gas be heated in contact with metallic potassium, calcium, or magnesium, the metals are oxidized, and the carbon of the gas separates in a solid condition. $\text{Ca} + \text{CO}_2 = 2\text{CaO} + \text{C}$, the calcium oxide reacts with the excess of carbon dioxide forming calcium carbonate: $\text{CaO} + \text{CO}_2 = \text{CaCO}_3$; while the carbon itself reduces the carbon dioxide: $\text{C} + \text{CO}_2 = 2\text{CO}$. Pass carbon dioxide from a suitable apparatus, A, Fig. 250, through a washing bottle, B, containing sulphuric

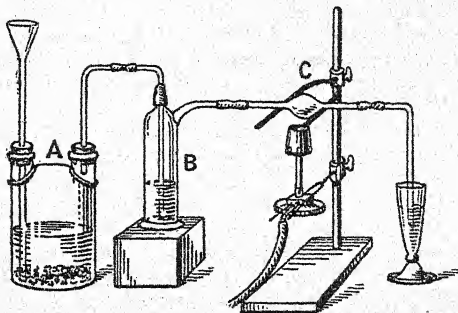


FIG. 250.—The Action of Metals on Carbon Dioxide.

acid, and then into a bulb of hard glass, *C*, containing a few shavings of metallic magnesium, potassium, or calcium. When all the air has been expelled from the apparatus, the bulb is heated. The carbon dioxide is decomposed vigorously, forming black carbon and calcium carbonate: $\text{CO}_2 + 2\text{Ca} \rightarrow 2\text{CaO} + \text{C}$; and $\text{CaO} + \text{CO}_2 = \text{CaCO}_3$. Metals like iron and zinc give carbon monoxide instead of carbon when heated with carbon dioxide: $\text{CO}_2 + \text{Zn} \rightleftharpoons \text{ZnO} + \text{CO}$.

If a rapid stream of carbon dioxide be passed through pure water in which rods of amalgamated zinc are immersed, a little of the carbon dioxide is reduced to formaldehyde, H.CO.H , thus: $\text{CO}_2 + 2\text{H}_2 = \text{H}_2\text{O} + \text{H.CO.H}$. The yield is increased if a trace of ammonia be present.

§ 3. Carbon Dioxide—Occurrence and History.

Occurrence.—Atmospheric air contains about 0.03 per cent. of its volume of carbon dioxide; and on account of the occurrence of this gas in air, T. Bergman (1774) called it "acid of air." It issues from the ground in many places both as a gas and in aqueous solution (mineral water). J. B. Boussingault (1844) estimated that Cotopaxi emitted more carbon dioxide per annum than was generated by life and combustion in a city like Paris. Owing to the fact that carbon dioxide is nearly one and a half times as heavy as air, this gas is inclined to collect as a gas in old deep wells, in valleys, and in depressions in the ground near lime kilns; and in certain neighbourhoods where carbon dioxide is evolved from volcanoes and fissures in the ground—*e.g.* the Valley of Death (Lake of Laach, Java), where one traveller says the whole of the bottom is strewn with the skeletons of human beings, animals, and birds which have been asphyxiated in an atmosphere overloaded with carbon dioxide. In the Grotta del Cane (Pozzuoli, Naples) there is said to be a constant depth of about 18 inches of carbon dioxide on the floor so that dogs entering the cave are suffocated while men are safe. An average of the analyses of the air in the Grotta del Cane, by C. St. C. Deville and F. le Blanc, gave 70.3 per cent. CO_2 ; 23.7 per cent. N_2 ; and 6.0 per cent. O_2 . It is said that the Emperor Tiberius had two slaves chained up in the cave, and that they died very quickly. The student must here distinguish between the *flow* of a heavy gas like carbon dioxide to the lowest possible level; and *diffusion* which leads to the dissipation of the gas into the atmosphere. If the supply of gas were not kept up, the gas which collects in the low levels would gradually be diffused through the atmosphere. The choke damp or black damp of mines and wells is air deficient in oxygen and rich in carbon dioxide. It is said that grizzly bears are sometimes found suffocated to death by carbon dioxide which issues from the ground in Death's Gulch (Yellowstone Park, Cal.). Tremendous deposits of calcium and magnesium carbonates—chalk, limestone, dolomite—as well as smaller deposits of other carbonates, occur in various parts of the world.

History.—Carbon dioxide has been known for a long time, but the early writers confused it with "air." J. B. van Jelmont (p. 23) called it *gas sylvestre*¹ to distinguish it from common air. He prepared it by the action

¹ The term "sylvestre" possibly refers to the supposed impossibility of condensing it; or it is to be taken literally as the "wild gas" which dwells in out-of-the-way places; or, having prepared the gas from charcoal, he thought of the *sylvan* surroundings of the charcoal burners.

of acids on alkalis and calcarerous substances; he showed that it was formed during the combustion of charcoal, during the fermentation and the decay of organic matter, and he recognized it in the "mineral water" at Spa (Belgium); in the Grotta del Cane (Naples), and other localities. Van Helmont also knew that the gas extinguished flame, and suffocated animals. Van Helmont, however, confused it with other gases which do not support combustion. J. Black (1755), however, proved the gas to be a peculiar constituent of carbonated alkalis, being "fixed" there in the solid state. Hence Black called the gas *fixed air*; and T. Bergman (1774), *aerial acid*. The chemical nature of carbon dioxide was clearly explained by Lavoisier who showed it to be an oxide of carbon.

§ 4. Liquid and Solid Carbon Dioxide.

Faraday liquefied carbon dioxide in 1823 by means of his sealed tube arrangement indicated in Fig. 109, and liquid carbon dioxide is now manufactured as a commercial article by pumping the gas into steel cylinders (bombs) by powerful compression pumps. The gas from, say, the fermenting vats of a brewery (p. 801) is washed, purified, and pumped into the bombs for "aërated water" manufacture, etc.

Properties of liquid carbon dioxide.—At -5° , carbon dioxide requires a pressure of 30.8 atmospheres for liquefaction; at $+5^{\circ}$, 40.4 atms.; and at $+15^{\circ}$, 52.1 atms.; and over 32° it cannot be liquefied by any known pressure. Liquid carbon dioxide is a colourless mobile liquid. It floats on water without mixing with it. It boils at -78.2° at atmospheric pressure. Carbon dioxide is used as a refrigerating agent on board ships where the use of ammonia for the same purpose is objectionable on account of the smell from slight leakages. According to a Board of Trade regulation, if ammonia refrigeration machines are used on board ship, the compression apparatus must have a special compartment.

Solidification of liquid carbon dioxide.—If liquid carbon dioxide be allowed to escape into the air from the nozzle of the bomb, the absorption of heat which attends the rapid evaporation causes a portion of the liquid to solidify. The solid is collected by tying a small canvas bag over the nozzle and inverting the bomb. By opening the nozzle for a few minutes, quite a lot of solid carbon dioxide can be collected. The solid can be shaken from the bag into a cardboard box for examination. It crystallizes in the cubic system.

Properties of solid carbon dioxide.—Solid carbon dioxide is a soft white snow-like substance—"carbonic acid snow." It evaporates in air without melting, but under a pressure of 5 atms. it can be melted to a liquid. It can be handled safely provided no pressure is applied. The effect of pressure is to break the film of gas between the solid and the warm hand and cause a severe burn, or rather, a blister, resembling the blister produced by a burn. A horn spoon can be used for handling the material. A piece of the solid, if placed inside an empty beaker, will evaporate sufficiently in a few minutes to fill the beaker with carbon dioxide gas to which the usual tests can be applied. Place some of the solid in a soda-water bottle and close the bottle with a rubber stopper. In a short time, the gas evolved will generate sufficient pressure to blow the stopper out of the bottle. Proper precautions should be taken in case the bottle should burst. A small beaker

can be placed on a few drops of water on a wooden block. Some solid carbon dioxide is placed in the beaker. In a few minutes, the beaker will be frozen to the wood.

Carbon dioxide "snow" dissolves in ether, and as the ether evaporates, a temperature approaching -110° can be obtained in air, and -140° under reduced pressure. The solution is a good conductor for heat, and serves as an excellent freezing mixture. A great many gases can be liquefied by passing them through tubes immersed in this mixture. Several interesting experiments can be made to illustrate the change in the properties of metals, etc., at low temperatures by means of solid carbon dioxide, or its solution in ether. Mercury freezes to a malleable solid resembling metallic lead.

The freezing of mercury is conveniently demonstrated as follows: A circular groove—1 cm. deep, 1 cm. wide, and 7 cm. diameter on the inside of the ring—

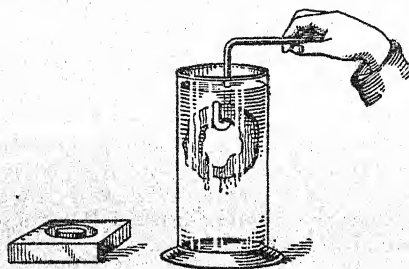


FIG. 251.—Freezing Mercury—F. G. Benedict.

a ring of ice on the hook while it runs to the bottom of the cylinder; the ice itself melts in a few minutes.

turned in a block of wood is filled about 7 mm. deep with dry mercury. Put the wooden mould in a large porcelain dish; fill the space around the mould with cotton wool; cover the mercury with a 7 mm. layer of solid carbon dioxide, or a mixture of carbon dioxide and ether. The mercury will be frozen in a few minutes. Invert the mould, and the mercury ring falls out. Lift the mercury ring rapidly with a glass hook, and lower it into a cylinder of water as illustrated in Fig. 252. The ring is at once covered with a layer of ice. The mercury melts before the ice, and leaves

§ 5. The Composition of Carbon Dioxide.

1. Composition by weight.—This has been established by the work of Dumas, Stas, Erdmann and Marchand, Roscoe, Friedel, etc. A weighed amount of carbon—diamond, graphite, etc.—is burnt at a red heat in a platinum boat, placed in a porcelain tube, *C*, Fig. 168. The tube also contains a layer of hot copper oxide. A stream of oxygen purified from carbon dioxide by passing through wash-bottles, *B*, containing potassium hydroxide and sulphuric acid, is led over the hot carbon. The resulting gas, on passing through the hot copper oxide, is all oxidized to carbon dioxide. The carbon dioxide is absorbed in weighed potash bulbs, *D*, and in a U-tube *E* containing soda lime in one leg, and calcium chloride in the other. The platinum boat containing the carbon is weighed before and after the experiment so that due allowance can be made for any ash present in the original sample of carbon. The following numbers are selected from the original list of experiments:

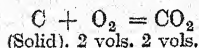
1.0000 gram of sugar charcoal required	2.6662 grams oxygen.
1.0000 gram of graphite required	2.6659 grams oxygen.
1.0000 gram of diamond required	2.6662 grams oxygen.

Hence the combining ratios of carbon and oxygen in carbon dioxide

are 8 grams of oxygen per 3.001 grams of carbon; giving 11.001 grams of carbon dioxide; or, 32 grams of oxygen require 12.004 grams of carbon to form 44.004 grams of carbon dioxide. The numbers obtained by different experimenters range from 11.99 to 12.00, and consequently 12 is taken to be the best representative value for the atomic weight of carbon (see p. 80). This experiment not only shows the relation between the weights of carbon and oxygen in carbon dioxide, but it also shows that the same relation obtains, within the limits of experimental error, whether the carbon be diamond, graphite, or charcoal. The chemical identity of the three last-named substances was established by S. Tennant (1796) and Mackenzie (1800) by showing that equal weights of each furnished the same weights of carbon dioxide.

2. Relative density of carbon dioxide.—This constant, determined by weighing an empty globe, and then the globe filled with gas, shows that if the density of oxygen is 32, that of carbon dioxide is 44.26. Hence the molecular weight of carbon dioxide is nearly 44.26. This is only possible if 12 parts of carbon are combined with 32 parts of oxygen by weight. With the atomic weight of oxygen 16, and carbon 12.003, it follows that the formula of carbon dioxide must be CO_2 .

3. Composition by volume.—The volume of a solid is negligibly small in comparison with the volume of the same substance in the gaseous state. Hence, if solid carbon be burnt in oxygen, Avogadro's hypothesis would lead to the inference that one volume of oxygen will form an equal volume of carbon dioxide:



This is best illustrated experimentally by means of the following modification of Hofmann's eudiometer charged with mercury, Fig. 252. The bulb of the right tube, Fig. 252, is charged with oxygen; and the stopper which carries a bone-ash crucible containing a chip of charcoal, is lowered into the position illustrated in the diagram. A slip of gummed paper is placed on the right tube indicating the position of the mercury when that liquid is at the same level in both tubes. The platinum wires are connected with an accumulator or battery. The small loop of platinum wire in contact with the carbon is thus heated red hot. This ignites the carbon which burns to carbon dioxide. The heat of the combustion expands the gas, but in a short time, when the apparatus has cooled, the level of the mercury is the same as before the experiment.¹ Hence carbon dioxide contains the equivalent of its own volume of oxygen.

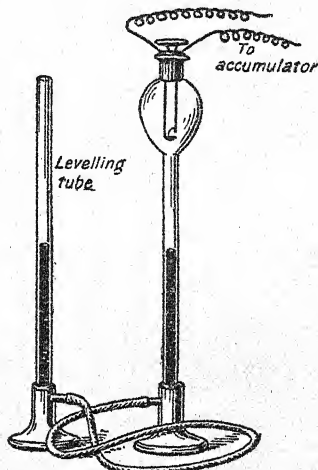


FIG. 252.—Volume Synthesis of Carbon Dioxide.

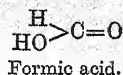
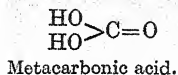
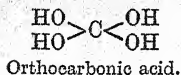
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¹ The necessary adjustment of the levelling tube during and after the combustion will be obvious.

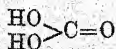
§ 6. Carbonic Acid and the Carbonates and Bicarbonates.

Carbonic acid.—Neither liquid nor dry gaseous carbon dioxide affects dry blue litmus paper, but if the paper be moistened, the litmus is coloured red, and the blue colour is restored when the paper is dried. There is little doubt that a small part of the carbon dioxide which dissolves in the water combines with the water to form carbonic acid. The aqueous solution of the gas turns blue litmus red.¹ Since carbon dioxide combines with water to form an acid, it is also called carbonic anhydride. The acid is probably $\text{O}=\text{C}(\text{OH})_2$ or H_2CO_3 . The acid is unstable, and has not been isolated. The acid decomposes at ordinary temperatures when the solutions are exposed to the air, and the carbon dioxide escapes. It is probable that we have to deal with the reversible reaction: $\text{H}_2\text{O} + \text{CO}_2 \rightleftharpoons \text{H}_2\text{CO}_3$. Only a very small quantity of the acid is formed as is evident from the low solubility of the gas in water. The system is in equilibrium when but a small proportion of the dissolved gas has produced H_2CO_3 . If a base be present, it will react with the carbonic acid and form a carbonate. More H_2O and CO_2 unite, and the resulting H_2CO_3 is removed by the base as fast as it is formed, until all the carbon dioxide in solution has been converted into carbonate.

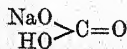
Orthocarbonic acid corresponding with $\text{C}(\text{OH})_4$, or H_4CO_4 , is not known, although some orthocarbonates, *e.g.* ethyl orthocarbonate, $\text{C}(\text{OC}_2\text{H}_5)_4$, are familiar to the student of organic chemistry. The ordinary "carbonates," now to be considered, are really metacarbonates derived from metacarbonic acid, $\text{CO}(\text{OH})_2$. The so-called formic acid may be regarded as a reduction derivative of metacarbonic acid made by replacing one hydroxyl group by hydrogen:



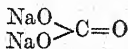
Alkali carbonates and bicarbonates.—As just indicated, the hydroxide $\text{C}(\text{OH})_4$ is unknown. The first dehydration product $\text{O}=\text{C}(\text{OH})_2$ is metacarbonic acid, or simply carbonic acid. Carbonates contain " CO_3 " as a dyad radicle. There are two possible alkali carbonates—normal and acid:



Carbonic acid.



Acid sodium carbonate.



Normal sodium carbonate.

The normal salt is formed during the first action of carbon dioxide on sodium hydroxide: $2\text{NaOH} + \text{CO}_2 = \text{Na}_2\text{CO}_3 + \text{H}_2\text{O}$. Sodium bicarbonate or sodium hydrogen carbonate or acid sodium carbonate, NaHCO_3 , is formed when an aqueous solution of the normal carbonate is treated with an excess of carbon dioxide: $\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightleftharpoons 2\text{NaHCO}_3$. Here the radicle " HCO_3 " acts as a monad. The acid alkali carbonates, RHCO_3 , are neutral to phenolphthalein, the normal carbonates are alkaline.

The great avidity of sodium hydroxide for carbon dioxide may be illustrated by several ingenious experiments. One of the simplest is to

¹ If the water contains traces of bicarbonates—*e.g.* hard water—the litmus will be coloured "wine-red" or "claret-red." Phenolphthalein is affected by carbonic acid as it is with other acids; methyl orange is not affected.

collect a cylinder of carbon dioxide over mercury and then pipette some sodium hydroxide solution under the edge of the cylinder. The sodium hydroxide rises to the top of the mercury, absorbs the gas, and the mercury rises in the cylinder accordingly. Potassium hydroxide is a better absorbent for carbon dioxide than the sodium hydroxide, because when a gas charged with carbon dioxide is bubbled through a wash-bottle the exit may become choked with solid sodium bicarbonate; because the sodium bicarbonate is less soluble in water than potassium bicarbonate (p. 810).

Solutions of potassium hydroxide are largely employed in analytical work as an absorbent for carbon dioxide. If air charged with carbon dioxide be drawn through a solution of sodium hydroxide in a weighed wash-bottle, the increase in weight represents the weight of carbon dioxide absorbed by the hydroxide. If the volume of the air be known, the amount of carbon dioxide in that volume of air follows at once. Similarly in gas analysis, the amount of carbon dioxide is determined from the contraction in volume which a given volume of the gas suffers after the gas has been in contact with sodium hydroxide.

The determination of carbon dioxide in a gas.—Hempel's burette, previously described, is fitted with a pipette like that shown in Fig. 254. The cylindrical part is filled with short cylindrical rolls of iron wire gauze 1 to 2 mm. mesh. The rolls are from 1 to 2 cm. long, and about 5 mm. thick. The cylinder is then closed with a rubber stopper, and the pipette is then charged with a 33 per cent. solution of potassium hydroxide in water. The iron does not oxidize during a determination because it is protected by the adherent solution, which thus helps to expose a large surface of the solvent to the action of the gas. The level of the potash solution is adjusted to a fixed point on the capillary tube, and connected with the gas burette by thick rubber tubing in the usual way. The measured volume of gas is driven from the burette to the pipette, and after standing for about a minute, the gas is returned to the burette. One absorption usually suffices to remove the carbon dioxide from a gas. The difference in the level of the burette before and after the absorption indicates the amount of carbon dioxide absorbed by the solution in the pipette.

In Pettenkofer's process (1858) for determining small quantities of carbon dioxide in air, an excess of a standard solution of barium hydroxide is introduced into a large glass cylinder, about 10 litres capacity, containing the air under investigation. The whole is shaken up. The carbon dioxide in the air is abstracted by the baryta water: $\text{Ba}(\text{OH})_2 + \text{CO}_2 = \text{H}_2\text{O} + \text{BaCO}_3$. The amount of barium hydroxide remaining in excess is then determined by titration with a standard solution of oxalic acid (using phenolphthalein as indicator). The amount of barium hydroxide

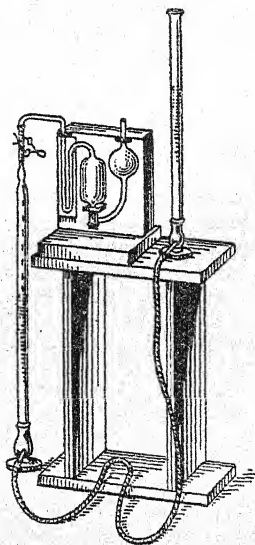


FIG. 253.—Analysis of Gases containing Carbon Dioxide.

in the original solution can be determined by titration. The *difference* represents the amount of barium hydroxide which has combined with the carbon dioxide in the air in the flask. The concentrations of the solutions are usually adjusted so that 1 c.c. of the baryta water will just precipitate 0.001 gram of carbon dioxide.

EXAMPLE.—50 c.c. of baryta water (1 c.c. = 1 mgrm. of CO_2) were shaken up with 10.8 litres of air at 7° and 765 mm. pressure, and the solution then required 42 c.c. of standard oxalic acid (1 c.c. = 1 mgrm. of the CO_2) for neutralization. Hence, 8 c.c. of the baryta solution were neutralized by the carbon dioxide in 10.8 litres of air, and consequently, 10,000 volumes of air contained 3.814 volumes of carbon dioxide.

The distinction of carbonates from bicarbonates or hydrocarbonates.—

(1) Normal carbonates give an immediate precipitate with magnesium sulphate; bicarbonates give a precipitate on boiling. (2) Calcium chloride may be used in place of magnesium sulphate. The boiling is to remove the carbonic acid which dissolves some of the carbonate. Filter, and either boil the clear solution or neutralize it with ammonia, when the carbonate is precipitated. 0.1 per cent. of sodium hydrocarbonate can be so detected in a normal carbonate. Ammonium salts interfere on account of the solubility of calcium carbonate in such solutions. (3) With mercuric chloride, normal carbonates give a reddish precipitate; bicarbonates give no precipitate. (4) Add a mixture of bleaching powder, potassium bromide, and chloroform to the solution; if hydrocarbonates be present, bromine is liberated and dissolved by the layer of chloroform.

Sodium carbonate crystallizes from water as a decahydrate: $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, in monoclinic prisms; potassium carbonate crystallizes as the anhydrous salt, K_2CO_3 . The former is efflorescent, the latter deliquescent. The solubilities of the carbonates and bicarbonates of lithium, sodium, and potassium—grams of salt per 100 c.c. of water at 20° —are as follows:

	Lithium.	Sodium.	Potassium.
Carbonate (R_2CO_3)	1.33	21.4	112.0
Acid carbonate (RHCO_2)	5.5	9.84	26.9

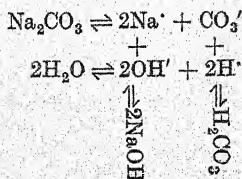
Sodium and potassium bicarbonates are thus less soluble than the normal salts. Lithium bicarbonate, like the bicarbonates of the alkaline earths, is more soluble than the normal carbonate. The rubidium and cesium hydrocarbonates are more soluble in alcohol, and also more stable than the potassium salt. They are converted into normal salts at 175° in an inert atmosphere. The solubility of sodium carbonate decreases with rise of temperature owing to the formation of lower hydrates. Thus between 30° and 50° , rhombic prisms of the heptahydrate, $\text{Na}_2\text{CO}_3 \cdot 7\text{H}_2\text{O}$, crystallize from the solution. Crystalline sodium carbonate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, melts at 60° , and on prolonged heating at this temperature, crystals of the dihydrate, $\text{Na}_2\text{CO}_3 \cdot 2\text{H}_2\text{O}$, are deposited. There is some doubt about the chemical individuality of this hydrate. The higher hydrates form the monohydrate, $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$, if dried over sulphuric acid; all the water is lost at 100° . Hot concentrated aqueous solutions of potassium carbonate at temperatures between 10° and 75° , deposit crystals of the trihydrate, $\text{K}_2\text{CO}_3 \cdot 3\text{H}_2\text{O}$; these crystals, at 100° , lose two molecules of water, forming the monohydrate: $\text{K}_2\text{CO}_3 \cdot \text{H}_2\text{O}$. The latter become anhydrous at 130° . Six hydrates of potassium carbonate have been reported, but the

evidence for those other than the two just mentioned is not considered satisfactory. Anhydrous sodium carbonate melts at about 850° , and potassium carbonate at 880° , mixtures of the two melt at lower temperatures. The eutectic mixture: $K_2CO_3 + Na_2CO_3$ —the so-called *fusion mixture*—melts about 690° .

The bicarbonates are transformed into the normal carbonates on calcination. Sodium bicarbonate is one ingredient of baking powder, and it is hence called "baking or cooking soda." When mixed with acid sodium tartrate (cream of tartar), which has a feeble acid reaction, carbon dioxide is evolved, this puffs up the dough, hence the term "*saleratus*" is sometimes applied to this salt—the salt which aerates, from the Latin *sal*, salt; *aer*, air or gas. Effervescent powders like "Seidlitz powder" are mixtures of sodium bicarbonate with tartaric acid, acid sodium tartrate, or some similar acid or salt. The mixture reacts when moistened, liberating carbon dioxide. Sodium bicarbonate is used in medicine for neutralizing the acidity of the stomach.

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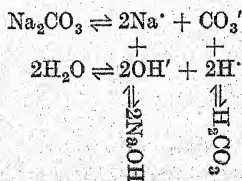
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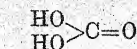
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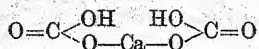


An aqueous solution of sodium carbonate behaves in many ways like a solution of sodium hydroxide—turns red litmus blue, etc.—and the ionic hypothesis thus shows how this action can be referred to the presence of an excess of OH' ions. If an hydrochloric acid solution be added to a solution of sodium carbonate, the H' ions of the acid unite directly with the OH' ions in the solution to form water, and if just sufficient HCl be added to remove the OH' ions, the solution will contain nothing more than Na' and Cl' ions such as would be obtained by dissolving sodium chloride in water. The CO_3'' ions of carbonic acid, and the Cl' ions of the hydrochloric acid, are competing for the Na' ions, but carbonic acid is very feeble in comparison with hydrochloric acid; and carbon dioxide formed by the *dissociation*, not ionization, of the carbonic acid is volatile under the conditions of the experiment and escapes from the solution.

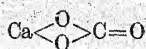
Calcium carbonate and bicarbonate.—Calcium carbonate has been previously studied, p. 390. Calcium carbonate is precipitated when a soluble carbonate, say, sodium carbonate, is added to a soluble calcium salt: $\text{CaCl}_2 + \text{Na}_2\text{CO}_3 = \text{CaCO}_3 + 2\text{NaCl}$. Calcium carbonate is also precipitated when a current of carbon dioxide is passed into lime-water: $\text{Ca}(\text{OH})_2 + \text{CO}_2 = \text{CaCO}_3 + \text{H}_2\text{O}$. Hence clear lime-water is rendered turbid by contact with carbon dioxide. This is a common test for carbon dioxide. Baryta water is still more sensitive. If an excess of carbon dioxide be passed through the solution, some of the calcium carbonate redissolves, and if the solution be not too concentrated, the turbidity may be clarified. This is due to the formation of a soluble calcium bicarbonate, $\text{Ca}(\text{HCO}_3)_2$ —acid calcium carbonate or calcium hydrogen carbonate: $\text{CaCO}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightleftharpoons \text{CaH}_2(\text{CO}_3)_2$. The relations of these carbonates will appear from the graphic formulæ:



Carbonic acid.

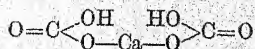


Acid calcium carbonate.

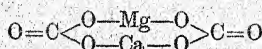


Normal calcium carbonate.

The acid carbonate is not very stable and it cannot be isolated by concentrating the aqueous solution, since it decomposes on exposure to the air owing to the escape of carbon dioxide and the precipitation of calcium carbonate; the solution likewise decomposes when boiled: $\text{Ca}(\text{HCO}_3)_2 = \text{CaCO}_3 + \text{H}_2\text{O} + \text{CO}_2$; and when treated with lime water, the two lime compounds react, precipitating calcium carbonate: $\text{Ca}(\text{HCO}_3)_2 + \text{Ca}(\text{OH})_2 = 2\text{CaCO}_3 + 2\text{H}_2\text{O}$. The bicarbonates of the alkaline earths have been isolated by mixing solutions of ammonium or potassium bicarbonate with the corresponding chloride all cooled to 0° . With calcium chloride, potassium bicarbonate furnishes a white crystalline precipitate with the ultimate composition $\text{Ca}(\text{HCO}_3)_2 : \text{CaCO}_3 \cdot 1.75\text{H}_2\text{CO}_3$. There is an interesting double carbonate of calcium and magnesium which occurs in nature as the mineral *dolomite*, and whose composition can be represented by the formula $\text{MgCO}_3 \cdot \text{CaCO}_3$, either as a solid solution, or as $\text{MgCa}(\text{CO}_3)_2$, a derivative of calcium hydrocarbonate:



Calcium hydrocarbonate.



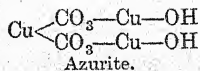
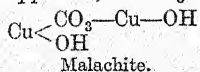
Dolomite.

The carbonates are not quite insoluble in water; 100 c.c. of water, in

contact with air, at 8.7° – 8.8° , dissolve 0.0016 gram of barium carbonate; 0.007 gram of calcium carbonate; and 0.00008 gram of strontium carbonate. Pure calcium carbonate imparts an alkaline reaction to water, probably owing to hydrolysis: $2\text{CaCO}_3 + 2\text{H}_2\text{O} \rightleftharpoons \text{Ca}(\text{HCO}_3)_2 + \text{Ca}(\text{OH})_2$. The precipitation of the carbonates of the alkaline earths by ammonium carbonate is a reversible reaction: $\text{CaCl}_2 + (\text{NH}_4)_2\text{CO}_3 \rightleftharpoons \text{CaCO}_3 + 2\text{NH}_4\text{Cl}$. Hence an excess of ammonium carbonate favours more complete precipitation, and the presence of ammonium chloride leads to a resolution of the precipitate. Magnesium carbonate is hydrolyzed by much water, and a basic carbonate is precipitated. A solution of magnesium carbonate in water saturated with carbon dioxide and containing about 13 grains of carbonate per fl. oz. is the *fluid magnesia* of the druggist. A basic magnesium carbonate is precipitated by adding sodium carbonate to the solution of a magnesium salt. The composition of the precipitate depends upon the temperature, concentration, etc., of the solutions at the time of precipitation. Cold dilute solutions of magnesium sulphate and sodium carbonate give the *magnesii carbonas levis* (light magnesium carbonate) of pharmacy, and hot concentrated solutions give *magnesii carbonas ponderosa* (heavy magnesium carbonate). The former is not so dense a powder as the latter; both are called *Magnesia alba*, and have approximately the same composition: $3\text{MgCO}_3 \cdot \text{Mg}(\text{OH})_2$. Some consider that the normal carbonate is first formed and immediately hydrolyzed by the water. The insolubility of magnesium hydroxide leads to its joint precipitation with the carbonate.

Cadmium, zinc, mercury, lead, and copper carbonate, like magnesium carbonate, are readily hydrolyzed, and produce basic carbonates. In the case of solutions of salts of iron and aluminium, alkaline carbonates precipitate the hydroxides. It is supposed that the hydrolysis of the aluminium carbonate is here complete. Ammonium carbonate gives a white precipitate of ferrous carbonate in a boiling solution of ferrous sulphate; the precipitate rapidly oxidizes, becoming green, bluish-green, and finally red ferric hydroxide. Air-free solutions should be employed, and the pharmacist's *ferri carbonas saccharata* is ferrous carbonate which has been rapidly mixed with a solution of sugar; this protects it from rapid oxidation. Normal zinc carbonate occurs native as *calamine*, ZnCO_3 , and the normal carbonate is precipitated from soluble zinc salts by potassium hydrogen carbonate. The excess of carbon dioxide is used to prevent hydrolysis. Cadmium does not form a normal carbonate.

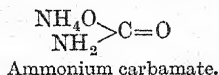
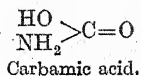
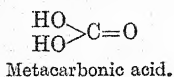
Two basic copper carbonates—*malachite*, $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$, and *azurite*, or *azure copper ore*, $2\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$ —occur in nature:



The former is usually green, the latter azure blue. Both have been made artificially by precipitation from solutions of copper salts by alkali carbonate. Crystals of malachite can be obtained by allowing a piece of chalk to stand in a solution of copper nitrate until basic copper nitrate crystals are formed, and allowing these to stand in a solution of sodium carbonate for a few days. Azurite is produced by heating crystallized copper nitrate with pieces of chalk under a pressure of about 4 atmospheres. *Verdigris* or copper rust has nearly the composition of malachite.

§ 7. Ammonium Carbonates.

In J. L. Gay-Lussac's study of the combination of gases by volume, he found that dried ammonia and carbon dioxide unite in the molecular proportions 2 : 1, forming a white solid. If the gases be very thoroughly dried no sign of chemical action occurs. It was at first thought that the white solid is ammonium carbonate, but it was afterwards found that whatever be the first product of the reaction a curious intramolecular change occurs and there is produced an ammonium salt of an amido-carbonic acid, now called carbamic acid, $\text{NH}_2\text{CO.OH}$, that is, ammonium carbamate, $\text{NH}_2\text{CO.ONH}_4$, or graphically :



The reaction is reversible $2\text{NH}_3 + \text{CO}_2 \rightleftharpoons \text{NH}_2\text{CO.ONH}_4$. Ammonium carbamate reacts with water forming normal ammonium carbonate, $(\text{NH}_4)_2\text{CO}_3$, symbolically $\text{NH}_2\text{CO.ONH}_4 + \text{H}_2\text{O} \rightleftharpoons (\text{NH}_4)_2\text{CO}_3$.

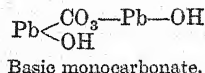
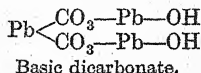
The alchemists obtained a *sal volatile* by heating animal matters in a closed vessel, or by distilling stale urine—*spiritus urine*. Extravagant recipes were sometimes given, and special medical virtues were attributed to *sal volatile* prepared in special ways. Thus, "5 lbs. of the skulls of persons who had been hanged should be distilled with 2 lbs. of dried vipers," and the product was extolled as a cure for epilepsy. R. Dossie showed in 1758 that *sal volatile* obtained in different ways was essentially the same product. The salt is made by subliming a mixture of two parts of calcium carbonate and one part of ammonium chloride or sulphate. The product when resublimed in the presence of a little water is a white fibrous mass of ammonium hydrocarbonate, $\text{NH}_4(\text{HCO}_3)$, mixed with a little ammonium carbonate which can be removed either by washing it out with alcohol, or exposing the salt to the air—when the carbamate volatilizes. Ammonium hydrocarbonate decomposes at 60° , and does not smell of ammonia; the commercial product smells of ammonia owing to the dissociation of the contained carbamate. The normal carbonate can be obtained by treating the commercial salt with a concentrated aqueous solution of ammonia at 12° . If the commercial salt is treated with concentrated ammonia at 30° , crystals of the so-called ammonium sesquicarbonate, $2\text{NH}_4(\text{HCO}_3).(\text{NH}_4)_2\text{CO}_3.\text{H}_2\text{O}$, are formed.

§ 8. Lead Carbonates—White Lead.

Normal lead carbonate, PbCO_3 , is precipitated when an alkaline bicarbonate is added to a soluble lead salt; a basic carbonate is precipitated if normal alkaline carbonate is used. It is also precipitated by adding ammonium carbonate to lead nitrate, or by passing carbon dioxide into a solution of lead nitrate or acetate. Native lead carbonate, PbCO_3 , is called *cerussite*. It is isomorphous with aragonite, CaCO_3 , and witherite, BaCO_3 . Barium does not form a basic carbonate, and normal barium carbonate does not lose carbon dioxide so readily as lead carbonate. The latter decomposes at about 200° into lead monoxide and carbon

dioxide; and at the same temperature it is readily reduced to the metal by carbon monoxide. The most important basic lead dicarbonate has approximately the composition $2\text{PbCO}_3 \cdot \text{Pb}(\text{OH})_2$; it is known as *white lead* or *ceruse*, and it was once made by exposing lead to the vapour of putrid urine. White lead is largely employed as a pigment, and in the manufacture of pottery glazes. It is made by a number of different processes.

The first product of the action of a slow current of carbon dioxide on a solution of dibasic lead acetate is a colloidal product, "insoluble" in cane sugar solutions, and with a composition corresponding with a basic lead monocarbonate, $\text{PbCO}_3 \cdot \text{Pb}(\text{OH})_2$. A precipitate corresponding with basic lead dicarbonate, $2\text{PbCO}_3 \cdot \text{Pb}(\text{OH})_2$, is obtained when the gas acts for a longer time; and the prolonged action of the gas gives a product approximating PbCO_3 . Some therefore argue that white lead must be a mixture of the basic carbonate $\text{PbCO}_3 \cdot \text{Pb}(\text{OH})_2$ with the normal carbonate, but in a more or less amorphous or colloidal state. Hence two hypotheses as to the nature of white lead are in the field: one postulates that white lead is a true chemical individual, a basic dicarbonate, or a basic monocarbonate. Graphically:



The other assumes that white lead is not a chemical individual, but rather an adsorption mixture, since both white lead and normal lead carbonate can withdraw lead hydroxide from solutions of basic lead acetate. The amount adsorbed—the equilibrium condition—is determined by the basicity of the solution. This argument was supported by E. Euston (1914) showing that calcium, barium, and zinc carbonates behaved similarly; thus, calcium carbonate forms an adsorption mixture, $2\text{CaCO}_3 \cdot \text{Pb}(\text{OH})_2$; barium sulphate, $3\text{BaSO}_4 \cdot \text{Pb}(\text{OH})_2$, etc. The fact that a solution of ammonium chloride will remove lead hydroxide from these mixtures, while a solution of sugar will not, is taken to mean that the carbonate and hydroxide in white lead are not merely mechanically mixed, but are yet so firmly united as to be properly considered in chemical combination.

Stack or Dutch process of manufacturing white lead.—The old so-called Dutch vinegar process—said to have been "old" in 1662—is supposed to give a better product than any other process. The time when ceruse was discovered is quite uncertain; both Pliny and Dioscorides speak of its preparation and of the use made of it and rouge, by the Roman ladies, as cosmetics to "improve" their appearance. A layer of 3 per cent. acetic acid is placed in the bottom of a large number of earthenware pots. A roll of thin metallic lead is placed on a shoulder in the pot above the acetic acid. The pots are placed upon a bed of spent tan bark from the leather-tanning yards.¹ The layer of pots is covered with boards. A layer of tan bark is placed on the boards, and then a second row of pots similarly charged. A chamber is nearly filled with pots in this manner. There are many variations in the methods of placing the pots, etc. For instance, in Cookson's works, Newcastle-on-Tyne, "straps" of "blue" lead are laid over rows of plain pots as illustrated in Fig. 254. The pots

¹ Wet hay or dung has also been used.

are charged with acid as before, and rested upon tan bark, and boards, resting on supports *A*, separate one layer from another. The lead acid pots and tan bark are thus confined in a series of little chambers. In about twelve weeks, most of the lead will have been transformed into compact masses of white lead. The stacks are then unloaded and the white lead is crushed in a mill; any lumps of unchanged lead are removed. The white lead is then ground up with water: passed into settling tanks, and finally dried.

The reactions which occur in each chamber are probably as follows: The heat developed by the decomposition of the tan bark volatilizes the

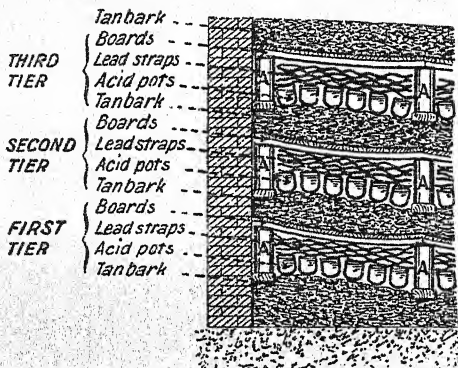


FIG. 254.—Section through Lower Corner of a White-lead Stack (Diagrammatic).

acetic acid which forms a basic lead acetate. For convenience in writing the equations, put \bar{A} , for the radicle of acetic acid and the acetates, *viz.* $C_2H_3O_2$. The reaction just mentioned is symbolized: $2H\bar{A} + 2Pb + O_2 = Pb\bar{A}_2 \cdot Pb(OH)_2$. The basic acetate so formed is decomposed by the carbon dioxide evolved during the decomposition of the tan bark. A mixture of normal lead acetate and basic lead carbonate is formed:

$3[Pb\bar{A}_2 \cdot Pb(OH)_2] + 2CO_2 = 3Pb\bar{A}_2 + 2PbCO_3 \cdot Pb(OH)_2 + 2H_2O$. The lead acetate in the presence of air and moisture reacts with more lead, forming more of the basic lead acetate: $2Pb\bar{A}_2 + 2Pb + O_2 + 2H_2O = 2[Pb\bar{A}_2 \cdot Pb(OH)_2]$. A small amount of lead acetate will thus serve for the manufacture of an indefinite amount of white lead.

The chamber corrosion process.—In this process straps of “blue” lead—about six inches wide and over a yard long—are hung on a series of horizontal bars arranged in tiers. Air, acetic acid vapour, steam, and carbon dioxide are admitted into the chamber maintained at the most favourable temperature. In about forty-five days the chambers are “drawn,” and the corroded lead is washed, etc., as before.

Many other more rapid processes have been proposed, *e.g.* electrolytic processes. In Thénard's process carbon dioxide is passed into a solution of basic lead acetate made by boiling litharge with lead acetate. The rapid processes may furnish a product of equal chemical purity with the stack processes, but the consumer frequently prefers white lead made by the older and more expensive stack process. The difference between white lead made by the different processes is not a question of chemical composition. The covering power of white lead made by the stack and corrosion processes is usually greater than that prepared by other processes. As a result less white lead is needed to impart to a given surface a given degree of opacity. White lead is liable to turn yellow or brown when exposed to the air of towns owing to the formation of lead sulphide, PbS . For some

purposes, therefore, white paints made from zinc oxide, lead sulphate, or barium sulphate are preferred in spite of their low covering power.

§ 9. Hard and Soft Water.

Water comes very near to the "universal solvent" (p. 189) of the alchemists' dreams. Natural water holds carbon dioxide in solution, and when such water comes in contact with magnesium and limestone rocks, some may be dissolved. Water holding magnesium and calcium salts in solution is said to be a **hard water**. The term is applied on account of the difficulty of obtaining a soap lather with such water. Soap is a compound of sodium with a fatty acid. The soap is decomposed by magnesium or calcium salts. The fatty acid unites with the latter to form an insoluble curdy precipitate. This action continues until all the lime and magnesian salts have been precipitated. Any further addition of soap at once produces a lather, and the soap can then be used as a cleansing agent. If a solution of soap of definite strength and a definite volume of water be employed—and this can be made by finding what volume of a given soap solution is required to produce a permanent lather with a solution of calcium chloride of known concentration—say equivalent to one gram of calcium carbonate per litre—the hardness of a given sample of water can be represented in terms of the amount of soap required to produce a lather—T. Clark's soap test. The hardness of water thus refers to the "soap-destroying power" of the water, and it is expressed in degrees. *Each degree of hardness corresponds with one grain of calcium carbonate, or its equivalent in other calcium or magnesium salts, per gallon of water.*

EXAMPLE.—6 c.c. of a standard soap solution (1 c.c. = 0.001 grm. CaCO_3) were required in titrating 50 c.c. of water in order to produce a lather which persisted for 5 minutes after the bottle containing lather and soap solution had been well shaken. Hence, the water had 12 parts of calcium and magnesium salts—hydrocarbonates, sulphates, nitrates, and chlorides—equivalent to 6.4 grams of calcium carbonate by weight per 100,000 c.c. of water. To convert this number into grains per gallon, multiply by 0.7. The result, 8.4, represents the number of grains of calcium carbonate per gallon of water. The hardness of the water on Clark's scale is therefore 8.4.

Hardness is also expressed in parts of calcium carbonate, or its equivalent, per 100,000 parts of water. Waters containing but small quantities of lime and magnesian salts lather freely with soap and they are accordingly called **soft waters**. A water less than 5° hardness may be called "soft," and a water between 18° and 20° hardness is "moderately hard," and if over 30° hardness, "very hard." Very soft waters are liable to attack metals like lead, zinc, and iron.

The process of removing lime salts from hard water is called **softening the water**. In the case just cited, soap is the softening agent. If the hardness of the water be due to the presence of acid carbonates of calcium or magnesium, mere boiling will soften the water because, as indicated above, the acid carbonates are then decomposed, and the normal carbonates are precipitated. In T. Clark's process for softening water (1841), the necessary amount of milk of lime or lime-water is added to convert all the acid carbonates of lime and magnesium into the normal carbonate. We have here the curious paradox—"add lime to remove lime." The theory of the action has been previously discussed, p. 813. Lime containing magnesia is best avoided in water softening, since the magnesia does no

effective work in this connection, and merely increases the amount of sludge to be handled.

Hard water may be wholly or partially softened by boiling. That part of the hardness so removed is termed **temporary hardness**; the degree of hardness which remains after prolonged boiling is called **permanent hardness**. Temporary hardness is due to the presence of calcium and magnesium bicarbonates; permanent hardness is due to the presence of calcium and magnesium chlorides or sulphates.¹ After removing the temporary hardness, if present, by boiling or by Clark's lime process, the permanent hardness may be removed by the addition of sodium carbonate. The boiling as well as the alkali carbonate processes of softening water were known towards the middle of the 18th century and were described by T. Bergman (1778). The sodium carbonate precipitates the calcium and magnesium as insoluble carbonates: $\text{CaSO}_4 + \text{Na}_2\text{CO}_3 = \text{CaCO}_3 + \text{Na}_2\text{SO}_4$. The water still contains sodium sulphate and sodium chloride, but the presence of a small quantity of these salts is not usually objectionable. Sodium carbonate will remove temporary as well as permanent hardness. In R. Gans' zeolitic or permutite process (1907) the water is passed through a layer of powdered artificial zeolite—called commercially *permutite*—made by fusing a mixture of quartz sand, china clay, and soda ash. The calcium and magnesium salts in the water react with the zeolite forming insoluble calcium and magnesium zeolite, and a soluble sodium salt. This removes both temporary and permanent hardness. Iron and manganese salts are also removed by this process. The "exhausted" zeolite is revived by passing a concentrated solution of sodium chloride through it. The calcium and magnesium zeolite reforms sodium zeolite. The processes used for softening potable hard water may thus be summarized: (1) Distillation; (2) Boiling; (3) The addition of lime with or without soda ash or other chemicals; (4) The zeolite or permutite process. The first process is usually too expensive.

The Report of the Rivers Pollution Commission classified unpolluted water as shown in Table LII, which represents the average of a number of analyses of each type (parts per 100,000).

TABLE LII.—AVERAGE COMPOSITION OF UNPOLLUTED NATURAL WATERS.

Class of Water.	Total solids.	Organic carbon.	Organic nitrogen.	Ammonia.	Nitrogen (nitrates) and nitrites).	Total combined nitrogen.	Chlorine.	Hardness.		
								Temporary.	Permanent.	Total.
Rain	2.95	0.070	0.015	0.029	0.003	0.042	0.822	0.4	0.5	0.9
Upland surface. . .	9.87	0.322	0.032	0.002	0.009	0.042	1.130	1.5	4.3	5.8
Deep well . . .	43.78	0.061	0.018	0.012	0.495	0.522	5.11	15.8	9.2	25.0
Spring	28.20	0.056	0.013	0.001	0.383	0.396	2.49	11.0	7.5	18.5

¹ There is a kind of "pseudo-hardness" due to the insolubility of soap in water containing sodium chloride. The so-called "marine soaps" are fairly satisfactory for such waters.

Boiler scale.—The “furring” of kettles, and the formation of “boiler scale,” is due to the precipitation of calcium and magnesium salts. Boiler scale is a poor conductor of heat, and hence the efficiency of a boiler which has “scaled” is seriously impaired. The boiler scale and the metal have different rates of contraction and expansion by heat. If the water in the boiler gets low, and the metal overheated, the “scale” may separate from the metal. If cold water now runs into the boiler, the scale quickly cools, contracts, and cracks. Water pours through the cracks on to the hot metal; a large volume of steam is generated, and the sudden pressure may be great enough to burst the boiler. Hard water in steam boilers not only produces “boiler scale,” but it may corrode the boiler shell, and cause “foaming” and “bumping.” The corrosion and pitting of boilers is usually produced by soft waters from swampy districts which contain organic acids in solution; by water from mining districts containing mineral acids in solution (*e.g.* sulphuric acid from the oxidation of pyrites, *q.v.*); and by water containing magnesium or calcium chlorides and nitrates. A great many nostrums for preventing boiler scale and corrosion have been proposed. In some, the water is treated before it enters the boiler; in others, the water is softened in the boiler itself. Whatever agents be used, they must be cheap; easily applied; yield no acid when used; and precipitate the salts which make the water hard, in a flocculent powdery condition, easily blown from the boiler.

Among the commoner agents are: sodium or calcium hydroxide, and sodium carbonate discussed above; sodium aluminate, NaAlO_2 , which acts as represented by the equation: $2\text{NaAlO}_2 + \text{Ca}(\text{HCO}_3)_2 + 2\text{H}_2\text{O} = \text{Na}_2\text{CO}_3 + 2\text{Al}(\text{OH})_3 + \text{CaCO}_3$; barium aluminate, $\text{Ba}(\text{AlO}_2)_2$, acts in a similar way; sodium fluoride which acts: $2\text{NaF} + \text{Ca}(\text{HCO}_3)_2 = 2\text{NaHCO}_3 + \text{CaF}_2$; normal sodium phosphate which acts: $2\text{Na}_3\text{PO}_4 + 3\text{CaSO}_4 = 3\text{Na}_2\text{SO}_4 + \text{Ca}_3(\text{PO}_4)_2$; etc.

Hard water in nature.—Rain-water percolating through limestone rocks becomes highly charged with dissolved calcium bicarbonate; such water in dripping through the roof of a cave or subterranean cavern will be exposed to the air; as a result, some carbon dioxide escapes from the solution, and a certain amount of calcium carbonate is deposited. Each drop of water adds its own little share of calcium carbonate. The deposit grows—maybe on the roof, when it is called a stalactite; maybe on the floor, when it is called a stalagmite. All depends upon the time occupied by each drop in gathering and dropping, how fast the carbonate is deposited. Measurements of a stalactite in the Ingleborough Cavern (Yorkshire), made in 1839 and 1873, show that it grew at the rate of 0.3 inch per annum. The San Filippo spring (Tuscany) is said to deposit “lime” at the rate of 12 inches a month, and the spring has formed a bed of limestone rock 250 feet thick, $1\frac{1}{4}$ miles long, and $\frac{1}{3}$ mile wide. The stalagmite grows upwards from the ground, and the stalactite grows downwards, like an icicle, from the roof. In time, the two may meet and form a pillar. Fig. 255 conveys but little idea of the beauty of some limestone caverns in which stalactites and stalagmites have been growing. The photograph shows stalactites, stalagmites, and pillars which have no doubt been formed in this manner. There are some very fine grottos, caves, or caverns in Castleton and other parts of Derbyshire, in the Isle of Skye, Antiparos (Archipelago), in the Hartz (Germany), Auxille (France), New South Wales, Auckland, United States, and numerous other limestone districts.

The building stone called *travertine* (Tiberstone) is probably a limestone deposited from a mineral spring. The Colosseum and much of ancient and modern Rome were built with this stone.

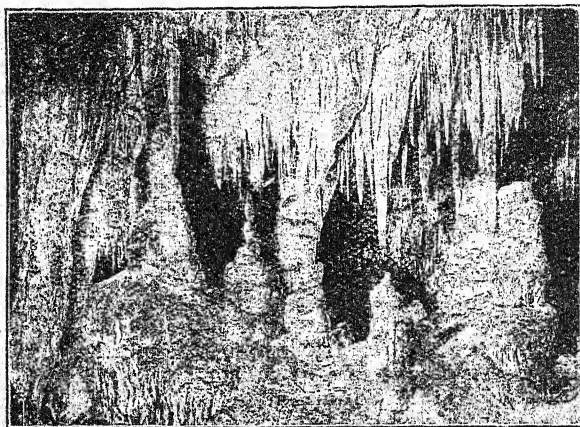


FIG. 255.—Stalactites and Stalagmites in the Yarrangobilly Caves, N.S.W.

§ 10. Alkali Manufacture.

The manufacture of soda is one of the greatest, if not the greatest of benefits which modern science has conferred on humanity.—A. ANATASI.

Before 1793, sodium carbonate was made from the ashes of seaweeds and potash carbonate from the ashes of land plants. Several methods were known at that time for converting sea salt into sodium carbonate, *e.g.* C. W. Scheele, about 1776, converted salt into soda by treating a solution of salt with lead oxide and passing carbon dioxide into the filtered solution. When Napoleon closed the European ports to English and American ships the main sources of supply of sodium carbonate and potashes were cut off. To ease the demand, Napoleon offered a prize of 100,000 francs for a process of manufacturing soda from common salt. The prize was won by N. Leblanc, 1794, who proposed a method of manufacture which held its own for over 100 years; although it has recently succumbed, for a long time it fought successfully against a serious rival in the ammonia-soda or Solvay's process; and both processes had to compete against electrolytic processes. Had it not been for the commercial value of the by-products of Leblanc's process, there can be little doubt that it would have been ousted long before it actually was.

As a matter of historical interest, Collison, in England, patented a similar process ten years earlier than N. Leblanc, *viz.* in 1782.

Leblanc's process is an interesting example of what could be done in application of a series of chemical operations on a large scale. The process was conducted in three stages: (i) The conversion of sodium chloride into sodium sulphate—the so-called *salt-cake*—by heating it with sulphuric

acid. (ii) The conversion of the salt-cake into *black-ash* by heating it to a high temperature with limestone and coke. Black ash is a mixture of calcium sulphide, sodium carbonate, etc.: $\text{Na}_2\text{SO}_4 + \text{CaCO}_3 + 2\text{C} = \text{CaS} + \text{Na}_2\text{CO}_3 + 2\text{CO}_2$. (iii) The extraction of sodium carbonate from the black-ash by leaching it with water, and purifying the product by crystallization. (iv) The recovery of sulphur from the undissolved black-ash. C. F. Claus and A. M. Chance suspended the residue in water, treated it with carbon dioxide: $\text{CO}_2 + \text{Ca}(\text{SH})_2 + \text{H}_2\text{O} = \text{CaCO}_3 + 2\text{H}_2\text{S}$. The hydrogen sulphide, mixed with sufficient oxygen to burn the hydrogen, but not the sulphur, was passed into a kiln containing iron oxide. By catalytic action, the iron oxide accelerated the oxidation of the hydrogen sulphide: $2\text{H}_2\text{S} + \text{O}_2 = 2\text{H}_2\text{O} + 2\text{S}$. Most of the sulphur which separated collected as a molten fluid at the bottom of the kiln or in an adjoining brickwork chamber, and was periodically run off. Tank waste was also used for the manufacture of sodium thiosulphate (*q.v.*). A factory in which soda ash is made is called an "alkali works." A works using the Leblanc process was divided into several departments: (1) The acid works where sulphuric acid was made; (2) Salt-cake works; (3) Black-ash works and lixiviation; (4) White-ash (soda ash) works; (5) Bleaching powder works where the hydrogen chloride from the salt-cake works was converted into chlorine and the latter converted into bleaching powder; and (6) Sulphur extraction from tank waste.

J. Thomsen's cryolite process.—A small proportion of the world's supply of sodium carbonate is made from cryolite by calcining a mixture of that mineral with calcium carbonate: $\text{Na}_3\text{AlF}_6 + 3\text{CaCO}_3 = \text{Na}_3\text{AlO}_3 + 3\text{CaF}_2 + 3\text{CO}_2$. The product—called *cryolite ash* is leached with water. Sodium aluminate passes into solution. This is treated with carbon dioxide: $2\text{Na}_2\text{AlO}_3 + 3\text{CO}_2 + 2\text{H}_2\text{O} = 2\text{Al}(\text{OH})_3 + 3\text{Na}_2\text{CO}_3$. The precipitated alumina is mainly used for making alum; the sodium carbonate is crystallized from the solution.

Electrolytic processes.—The electrolytic processes of manufacture depend upon the formation of sodium hydroxide during the electrolysis of aqueous solutions of sodium chloride. The chlorine obtained as a by-product is used in the manufacture of bleaching powder. The hydroxide is converted into sodium carbonate by treatment with carbon dioxide, which is obtained as a by-product in the fermentation industries, etc. Everything is utilized so that the cost essentially depends upon the price of the current used.

§ 11. The Ammonia-Soda or E. Solvay's Process.

It is claimed that this process was used by Fresnel in 1811, by Vogel in 1822, and by J. Thom in 1836; but the real history of the process dates from the patent of H. G. Dyer and J. Hemming in 1838. Several improvements were subsequently devised by numerous chemists, but the process was not successful until E. Solvay took it in hand between 1861 and 1872. This process depends upon the fact that when a concentrated solution of sodium chloride is saturated with ammonia, and carbon dioxide is passed through the mixture, ammonium carbonate is formed in the first stage of the operation, and that this reacts with the sodium chloride producing

sodium carbonate: $2\text{NaCl} + (\text{NH}_4)_2\text{CO}_3 \rightleftharpoons \text{Na}_2\text{CO}_3 + 2\text{NH}_4\text{Cl}$. Sometimes a little sodium carbonate precipitates before all the ammonia is carbonated, in any case some sodium carbonate is precipitated and that remaining in solution is further carbonated: $\text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} = 2\text{NaHCO}_3$, and then precipitated. It is highly improbable that the ammonia is converted into ammonium bicarbonate, $(\text{NH}_4)\text{HCO}_3$, which then undergoes double decomposition with the sodium chloride: (a) because ammonium bicarbonate is just as sparingly soluble as sodium bicarbonate in the mother-liquids, and would therefore be precipitated; and (b) such a reaction would be endothermal, whereas the actual reaction is exothermal in agreement with: $\text{CO}_2 + \text{H}_2\text{O} + \text{Na}_2\text{CO}_3 = 2\text{NaHCO}_3 + 18.4 \text{ Cals.}$ The reactions are all reversible, and the amount of sodium chloride converted into the bicarbonate depends on the relative concentrations of all four salts. W. Mason's analysis (1914) of the solution when the precipitation begins shows that 50 per cent. of the sodium chloride has decomposed and 11.8 per cent. of sodium carbonate is in solution. The mother liquid remaining after the precipitation of the bicarbonate contains in grams per 100 c.c.:

NaCl	$(\text{NH}_4)_2\text{CO}_3$	Na_2CO_3	NH_4Cl
8.5	8.1	0.5	78.1

while the precipitate contains:

NaCl	NaHCO_3	Na_2CO_3	NH_3	H_2O	per cent.
0.2	71.3	4.0	0.5	24.0	

The aim of the industrial chemist is to establish the conditions necessary to make the left to right reaction almost complete, a certain fraction of sodium chloride is certain to escape conversion. If sodium carbonate is needed the bicarbonate is calcined: $2\text{NaHCO}_3 = \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$. This carbon dioxide forms part of that used in the first state of the operation. The mother liquid remaining after the separation of sodium bicarbonate is treated with lime obtained by burning limestone: $\text{CaCO}_3 = \text{CaO} + \text{CO}_2$; and the ammonia is recovered: $2\text{NH}_4\text{Cl} + \text{CaO} = \text{CaCl}_2 + \text{H}_2\text{O} + 2\text{NH}_3$. The ammonia and carbon dioxide evolved in these two operations are used again. Thus calcium chloride is the only by-product which is not utilized. Great care has to be exercised in the manufacturing process to avoid losses of the relatively expensive ammonia; the chlorine of the salt appears as a by-product—calcium chloride—in a form which cannot be economically recovered. Actually more salt is needed in the manufacturing process for a given weight of carbonate than in Leblanc's process. On the other hand, the ammonia-soda process shows its superiority to Leblanc's process: (1) The manufacturing costs are less because brine direct from the salt beds can be used; (2) There are no noxious by-products; and (3) The final product is less contaminated with impurities.

The process can be illustrated by connecting an apparatus, A, for generating ammonia, Fig. 256, and an apparatus, B, for making washed carbon dioxide with a tower, C, filled with a saturated solution of sodium chloride and fitted with four perforated iron discs as shown in the diagram. The tower is provided with an exit tube dipping in a beaker of water. The solution is first saturated with ammonia, and then with carbon dioxide. In about an hour, crystals of sodium bicarbonate will be deposited on the perforated shelves

Potassium carbonate cannot be made by Solvay's process because acid potassium carbonate is too soluble. A great deal of the potassium carbonate of commerce is made by Leblanc's process. It is also made from the potassium chloride or sulphate of the Stassfurt deposits by forcing carbon dioxide into a solution of potassium chloride containing magnesium chloride. A complex double compound of magnesium carbonate and potassium hydrogen carbonate is formed. This is decomposed by steam; the magnesium is precipitated as hydroxide; and the clear solution is evaporated for potassium carbonate.

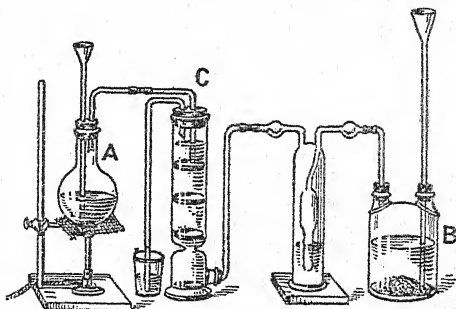
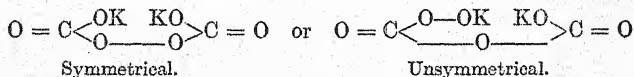


FIG. 256.—Illustration of Solvay's Process.

§ 12. Percarbonates.

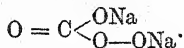
E. J. Constan and A. von Hansen (1896) found that when a saturated aqueous solution of potassium carbonate is electrolyzed at a temperature below -10° , a sky blue precipitate separates, which, when dried over phosphorus pentoxide, is nearly white. The precipitate is a mixture of potassium carbonate and potassium percarbonate— $K_2C_2O_6$ or KCO_3 , which may be graphically represented by one of the formulæ:



Neither the acid $H_2C_2O_6$, nor the anhydride C_2O_5 has been isolated. The percarbonate decomposes on gentle warming, giving potassium carbonate and a mixture of carbon dioxide and oxygen: $2K_2C_2O_6 = 2K_2CO_3 + 2CO_2 + O_2$. It furnishes hydrogen peroxide when treated with dilute sulphuric acid or dilute alkalis: $K_2C_2O_6 + 2H_2SO_4 = 2KHSO_4 + 2CO_2 + H_2O_2$; $K_2C_2O_6 + 2KOH = 2K_2CO_3 + H_2O_2$. The salt thus acts as a strong oxidizing agent like hydrogen dioxide or sodium dioxide. For instance, it converts lead sulphide (PbS) into lead sulphate ($PbSO_4$), bleaches indigo, etc. It forms with manganese and lead dioxides the corresponding carbonates, and at the same time reduces the dioxide: $PbO_2 + K_2C_2O_6 = PbCO_3 + K_2CO_3 + O_2$. The dry salt keeps unchanged for an indefinite period, and some accordingly prefer potassium percarbonate to less stable salts as a source of hydrogen peroxide in certain analytical operations: oxidation of sulphides to sulphates, chromic salts to chromates, hypochlorites to chlorides, titanium and vanadium to the coloured peroxides, etc. A similar compound, isomeric with $K_2C_2O_6$ prepared by electrolysis, can be made by the action of carbon dioxide on a cold solution of sodium peroxide. This compound behaves differently towards a solution of

potassium iodide whether added as solid or as a solution, and differently from the percarbonate prepared by electrolysis.

S. Tanatar (1899) prepared yet a third percarbonate by dissolving an alkaline carbonate in a 3 per cent. solution of hydrogen peroxide, and after a few minutes, precipitating the salt by the addition of ether. The same salt is obtained by the action of carbon dioxide on a solution of, say, sodium peroxide in water. The salt has the ultimate composition Na_2CO_4 . Just as the acid H_2CO_3 corresponds with sulphurous acid H_2SO_3 , so the salt Na_2CO_4 might be supposed, at first sight, to correspond with salts of sulphuric acid— H_2SO_4 ; but the constitution of Tanatar's percarbonate is generally taken to be



Salts of the type Na_2CO_4 are called *monoperoxycarbonates*—e.g. $(\text{NH}_4)_2\text{CO}_4 \cdot 2\text{H}_2\text{O}$ —and salts of the type $\text{Na}_2\text{C}_2\text{O}_6$ are called *monoperoxydicarbonates*. The acid corresponding with H_2CO_4 has not been prepared. The aqueous solutions of the salt Na_2CO_4 readily decompose into sodium carbonate and hydrogen peroxide. On heating, the salt forms sodium carbonate and oxygen, but no carbon dioxide.

The alkaline solutions decompose rapidly while the acidified solution (presumably containing hydrogen dioxide) is much more stable.

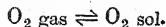
Some of the so-called percarbonates prepared by the action of hydrogen peroxide on the carbonates seem to contain "hydrogen peroxide of crystallization"—e.g. $\text{Rb}_2\text{CO}_3 \cdot 3\text{H}_2\text{O}_2$ —and these salts liberate very little iodine from a neutral solution of potassium iodide under conditions where the true percarbonates give a quantitative yield of iodine. It may be added, *en passant*, that similar remarks apply to the sulphates and persulphates. Hence, a distinction must be drawn between the true peroxy-salts, and salts containing hydrogen peroxide of crystallization.

§ 13. The Solubility of Gases in Liquids.

Our study of the solubility of gases in liquids (p. 188) may now be resumed. The solubility depends upon the nature of the gas and of the solvent, as well as upon the temperature and pressure of the system. No "common" solubility has been observed, and we have no generalization of such wide applicability as Boyle's and Charles' laws. Some gases are only slightly soluble in water, others dissolve very copiously. One volume of water, at normal temperature and pressure, will dissolve 1200 volumes of ammonia, and but 0.021 volume of hydrogen.

Kinetic theory of gaseous solution.—The kinetic theory of gases furnishes a mental picture of the process of solution of a gas in water. Suppose that a gas-free liquid be brought into a vessel containing a gas. The molecules of gas impinging upon the surface of the liquid will be absorbed. The dissolved molecules move about in the liquid in all directions, a small number escape back into the gas above. As the molecules of the gas crowd more and more in the liquid, the number of molecules absorbed by the liquid becomes more and more nearly equal to the number which escapes back into the superincumbent gas. If the pressure of the gas remains constant, a time will come when the number of gas molecules which leave the liquid will be equal to the number absorbed.

The system is then in a state of dynamic equilibrium resembling the equilibrium of a vapour in contact with its own liquid, and the solution is saturated with the gas under the given conditions of temperature and pressure.



The surface of the liquid in contact with a dissolving gas must be very quickly saturated with the gas, and the rate of absorption of a gas by a liquid at rest is really a measure of the rate of diffusion of the gas from the surface through the body of the liquid. The molecules of the liquid must have some attractive influence on the molecules of the dissolved gas.¹

Influence of pressure.—With gases which are not very soluble in liquids, the greater the pressure, the more soluble the gas; that is, the greater the pressure, the greater the concentration of the gas in the solution. W. Henry (1803) discovered an important relation between the pressure and the solubility of a gas. Under a total pressure, p , of 2 atmospheres, a saturated solution of gas holds in solution twice as much gas by weight, w , as under a total pressure of 1 atmosphere. Otherwise expressed,

the weight, w , of gas absorbed by a given volume of liquid is directly proportional to the pressure, p , or $w/p = \text{a constant}$. According to Boyle's law, the concentration of a gas, or the amount of gas w in an enclosed space, is proportional to the pressure. One volume of a gas at atmospheric pressure will make half a volume

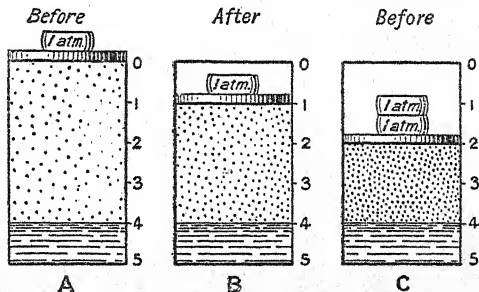


Fig. 257.—Diagrammatic Illustration of Henry's Law.

of gas at a pressure of 2 atmospheres. Under a pressure of 4 atmospheres, a saturated solution of gas holds four times as much gas in solution as it did under a pressure of 1 atmosphere, but four volumes of gas at atmospheric pressure occupy but one volume at a pressure of 4 atmospheres. Imagine a cylinder graduated into five equal divisions, Fig. 257, and one division filled with water, and four divisions with carbon dioxide. The cylinder is closed with a lightly fitting frictionless piston. When everything is in equilibrium, the water dissolves its own volume of carbon dioxide and the piston falls one division. Imagine the gas now compressed 2 atmospheres; the piston will be at division 2 and as the liquid dissolves its own volume of gas, the piston falls to the 3 division. Hence follows Henry's law: "under equal circumstances of temperature, water takes up in all cases the same volume (c) of a condensed gas as of gas under ordinary pressure." That is to say, $c/p = \text{a constant}$. The law thus describes the behaviour of the less soluble gases very well—carbon monoxide, nitrogen, hydrogen,

¹ It is probable that the attractive forces between the molecules of the liquid and the dissolving gas determine the coefficient of solubility of the different gases, otherwise we should expect the solubility of certain groups of gases to be the same.

oxygen—but not the more soluble gases like ammonia, hydrogen chloride, sulphur dioxide. The deviation is not very great with carbon dioxide, though it is appreciable :

Pressure (<i>p</i>) . . .	1	5	10	15	20	25	30	atms.
Solubility (<i>c</i>) . . .	1.80	8.65	16.03	21.95	26.65	30.55	33.74	vols. per c.c.
<i>c/p</i>	1.80	1.73	1.60	1.46	1.33	1.22	1.12	

The graph, Fig. 258, represents the observed values of *p* and *c*; and the dotted curve in the same diagram represents what the graph would have been had carbon dioxide behaved as described by Henry's law. The value of *c/p* is not therefore quite constant. The constant for the solubility of ozone in water is $\frac{1}{10}N\cdot H_2SO_4$, at 20° is 0.23, and at 0°, 0.44. In the case of carbon dioxide, the failure arises from the fact that carbon dioxide reacts chemically with water. Henry's law refers (1) to gases which do not act chemically on the solvent. When carbon dioxide dissolves in water, one portion enters into combination to produce a new substance—carbonic acid—while the other portion dissolves in the physical sense as carbon dioxide. The latter portion alone comes within the

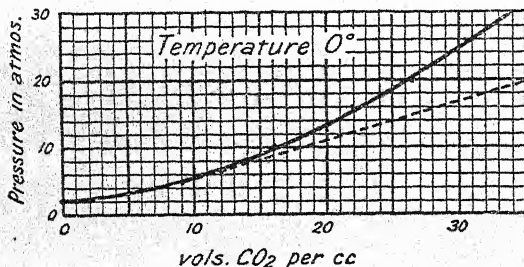


FIG. 258.—Solubility Curve of CO_2 in Water.

province of Henry's law. The condition of the carbon dioxide dissolved by the water is therefore represented by: $CO_2 + H_2O \rightleftharpoons H_2CO_3$.

Henry's law also assumes (2) that the molecules of the dissolved gas are no heavier than the molecules of the gas itself. Hence it may

be inferred that if a gas "obeys" Henry's law, it will have the same molecular weight in solution and in the gaseous condition. Henry's law is therefore to be regarded as a link connecting the molecular weights of gaseous and dissolved substances with one unit of measurement. If 32 grams of oxygen depress the freezing point to the same extent as 342.2 grams of cane sugar, it is inferred that the molecular weights of oxygen and cane sugar are related as 32 : 342.2; and since oxygen gas has a molecular weight of 32, it is assumed that cane sugar if it could be vaporized, and if its gas obeyed Henry's law, would have a molecular weight of 342.2, because, as indicated above, the molecular weights of a substance in solution and in the gaseous state are the same.

Henry's law and the kinetic theory.—The gas is in equilibrium with its own solution when the number of molecules which escape from the solution is the same as those which are captured by the solution in a given time. By, say, doubling the pressure the molecular concentration will be doubled, the gas molecules will be crowded more closely together, and the rate at which the solution captures the molecules will be increased twofold for the new state of equilibrium. Similarly the rate of escape will be doubled, and consequently the number of dissolved molecules will

be doubled. Hence variations of pressure do not alter the relative number of molecules per unit volume of solution and of gas; and the volume of gas dissolved will be independent of the pressure on the gas, while the weight of gas dissolved will be directly proportional to the pressure.

Measuring the solubility of gases.—The solubility of gases in a liquid can be conveniently determined in the following apparatus, modified from that used by R. Heidenhain and L. Meyer (1863). The pipette *C* (Fig. 259) is filled with a measured volume of the liquid under investigation. It is connected with a Hempel's burette by means of a piece of metal tubing of narrow bore. The gas under investigation is introduced into the measuring tube *A*, *via* the three-way cocks *C* and *A*, by first raising and then lowering the levelling tube *B*. A definite volume of liquid is then run from the pipette *C* by opening the lower cock and putting the pipette *C* in communication with the burette. A certain amount of gas enters the pipette. The contents of the pipette are then agitated, and when all is in equilibrium, and the liquid is saturated and the mercury in the burette and levelling tube are at a constant level, place the lower end of the absorption pipette beneath a vessel of mercury, and bring the liquid in the pipette to its former level. The diminution of the volume of gas in the burette represents the volume of gas absorbed by the volume of liquid in the pipette at the temperature and barometric pressure at the time of the experiment. To vary the temperature of absorption, the burette and pipette must be kept in a liquid or vapour bath at the desired temperature, and the measurements made when everything is in equilibrium.

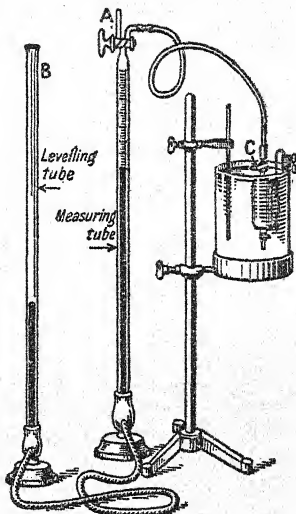


FIG. 259.—Determination of the Solubility of Gases.

R. Bunsen (1855) applied the term absorption coefficient to the volume of gas reduced to n.p.t. which one volume of the liquid will absorb at θ° when the pressure is 760 mm. Thus, the coefficient of absorption of hydrogen at 7° is 0.020, meaning that at this temperature one volume of water will absorb a volume of hydrogen equivalent to 0.020 volume at n.p.t. Thus, if v volumes of gas are absorbed by u volumes of liquid at T and p mm. pressure, then if $vpT_0/760T$ represents the volume V reduced to n.p.t. it follows from Henry's law that if this volume be absorbed at a pressure p , $760/p$ times this volume will be absorbed at 760 mm. pressure. Hence, the volume of gas absorbed by u volumes of liquid at n.p.t. will be vT_0/T , or the amount dissolved by unit volume of liquid will be vT_0/uT , and this is the coefficient of absorption of the gas. W. Ostwald (1888) used the term solubility for the ratio of the volume of gas absorbed to the volume of the absorbing liquid at any specified temperature and pressure. Thus, for the above-mentioned condition, the solubility of the gas at T° is v/u , that is, T_0/T times the coefficient of absorption.

EXAMPLE.—O. F. Tower (1906) observed that 27.22 c.c. of 98 per cent. sulphuric acid at 18.5° and 714.6 mm. barometric pressure absorbed 0.53 c.c. of nitrogen. Hence, the solubility coefficient of the gas in this acid is $0.58/27.22 = 0.0213$, or reduced to 760 mm. 0.0227.

It might be emphasized, in passing, that when a gas is dissolved in a liquid at a given temperature, the ratio between the concentration of the gas in the liquid and in the space above is always the same. Thus Henry's law is a law of distribution for gases because it describes the way a gas distributes itself between the solvent and the space above. Henry's law also describes the condition of equilibrium of a gas whose molecules are physically and chemically independent of each other, and of the solvent.

The influence of temperature.—The solubility of a gas in a liquid is very sensitive to changes of temperature. The higher the temperature, the less the solubility of the gas. The solubility curve for helium is not much affected by changes of temperature up to 50° , but what little effect there is seems to indicate that the solubility of the gas *increases* as the temperature rises from 25° to 50° . Hydrogen was once supposed to behave in a similar way, between 0° and 25° , but later, more careful measurements show that the solubility decreases steadily from 0.0214 at 0° to 0.0171 at 26° .

§ 14. The Solubility of Mixed Gases—Dalton's Law.

When a mixture of two gases is exposed to the action of a solvent, the quantity of each gas dissolved by the liquid depends upon the amount and the solubility of each gas present. The amount of each gas determines its partial pressure, and since the partial pressure of each gas is independent of the others, it follows that when a mixture of gases is exposed to the action of a solvent, the amount of each gas which is dissolved by the solvent is proportional to its partial pressure. Each gas behaves as if the others were absent. This is Dalton's law. It is a simple extension of Henry's law.

Application to air.—When air containing, say, 79 volumes of nitrogen¹ and 21 volumes of oxygen, and 0.04 volume of carbon dioxide, is shaken up with water, the amount of each gas absorbed by the water can be approximately computed in the following manner: The relative solubilities are: nitrogen, 0.02; oxygen, 0.04; and carbon dioxide, 1.79. The partial pressure of each gas is proportional to the relative amount of that gas present in a given volume of air. If the pressure of air be just one atmosphere, the partial pressure of the nitrogen will be proportional to 0.79×1 ; of oxygen, 0.21×1 ; and of carbon dioxide, 0.0004×1 . Hence the relative amounts of these gases absorbed by the water will be: nitrogen, $0.79 \times 0.02 = 0.0158$; oxygen, $0.21 \times 0.04 = 0.0082$; and carbon dioxide, $0.0004 \times 1.79 = 0.00072$. Hence 1 c.c. of water dissolves 0.0158 c.c. of nitrogen; 0.0082 c.c. of oxygen; and 0.00072 c.c. of carbon dioxide. The composition of the dissolved gases, if removed from the air by boiling, or exposure to a vacuum, will be: nitrogen, 63.9 per cent.; oxygen, 33.2 per cent.; carbon dioxide, 2.9 per cent. The relatively large solubility of carbon dioxide is counterbalanced by its low partial pressure, otherwise we might expect a heavy rain storm to remove a great part of the carbon dioxide from the atmosphere. Mallet has based a proposal to separate oxygen from atmospheric air on the action. If the carbon dioxide be removed by passing the air through

¹ Neglecting argon, etc.

an aqueous solution of sodium hydroxide, the oxygen and nitrogen in the remaining gases after the first absorption will be nearly in the proportion: nitrogen 65.7 per cent., and oxygen 34.3 per cent. If this mixture be driven from the water by boiling, and the mixture again treated with air-free water, a gaseous mixture containing 49 per cent. of oxygen is obtained; and after the eighth absorption, a gas containing 98 per cent. of oxygen results. The method is not practicable, though it is an interesting application of Henry's and Dalton's laws.

EXAMPLES.—(1) If a litre of water dissolves 25.12 c.c. of air, show that the mixture of gases obtained by boiling this water will contain 8.2 c.c. of oxygen; 15.8 c.c. of nitrogen; and 0.72 c.c. of carbon dioxide.

(2) Show that the mixture of gases expelled from water on boiling is between 71 and 72 times as rich in carbon dioxide as normal air.

(3) The solubility of hydrogen is 0.02, and of oxygen 0.04; calculate the volume of each of these gases which would be dissolved by 1000 c.c. of water from an electrolytic mixture of hydrogen and oxygen.

According to Wroblewsky, the crystalline hydrate, $\text{CO}_2 \cdot 8\text{H}_2\text{O}$, is formed when aqueous solutions of carbon dioxide at 0° are subjected to a pressure of 12 atmospheres; and P. Villard obtained crystals of the hydrate $\text{CO}_2 \cdot 6\text{H}_2\text{O}$; he likewise obtained crystals of the hydrates $\text{N}_2 \cdot 0.6\text{H}_2\text{O}$; $\text{SO}_2 \cdot 6\text{H}_2\text{O}$; and $\text{CH}_3\text{Cl} \cdot 6\text{H}_2\text{O}$.

§ 15. Van der Waals' attempt to obtain a more exact Gas Equation.

"All nature widens upwards. Evermore

The simpler essence lower lies.

More complex is more perfect, owning more

Discourse, more widely wise."—ANON.

We have seen that the gas equation does not describe the behaviour of real gases with respect to changes in volume with variations of temperature and pressure. The same gas does not behave in the same way at high and at low pressures. For many gases—hydrogen, oxygen, etc.—the laws of Boyle and of Charles are fairly exact at temperatures and pressures not far removed from normal atmospheric conditions; and

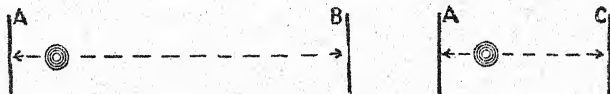


FIG. 260.

where they are not—carbon dioxide, ethylene, etc.—it is often convenient to neglect the small deviations. This means that gas calculations with $pv = RT$ are made upon imaginary gases, sometimes styled *ideal* or *perfect gases*. When the pressure upon the gas is very great, the error becomes quite appreciable, and it is necessary to revise the simple gas law: $pv = RT$.

(1) **Molecular co-volume.**—Let b denote the space occupied by the molecule as it moves to and fro between the boundary walls AB , Fig. 260. If this distance be halved, AC , Fig. 260, while the volume of the molecule remains constant, the molecule will have *less* than half its former distance

to pass from one side to the other.¹ It will therefore strike the walls more frequently than before. Hence the outward pressure of the molecule will increase more rapidly with decreasing volume than is described by Boyle's law. Boyle's law refers to the whole volume of the gas, but rather should it refer to the space in which the molecules move. We therefore write $v - b$ in place of v in Boyle's law, and the result is:

$$p(v - b) = RT$$

where b is called the "co-volume" or "vibratory volume" of the molecule.²

(2) Molecular attraction.—The closer proximity of the molecules to one another, the greater will be the effect of the attractive forces between the molecules. This attractive force will tend to make the gas occupy a smaller volume. The effect is much the same as if the gas were subjected to the action of a greater external pressure than the observed or apparent pressure of the gas. The assumption that the attractive force between the molecules varies inversely as the fourth power of the distance between the molecules leads to a/v^2 as the magnitude of the molecular attraction, where a is a constant which varies with the nature of the gas. Granting the assumption, a/v^2 must therefore be added to the observed pressure of the gas in order to indicate the total pressure tending to compress the gas. On correcting the equation $pv = RT$ for the volume and the cohesion of the molecules, we obtain the so-called J. D. van der Waals' equation (1872):

$$\left(p + \frac{a}{v^2}\right)(v - b) = RT$$

This amended equation agrees fairly well with a number of observations of gases under large pressures, and of gases which are near their points of liquefaction—*e.g.* ethylene, carbon dioxide, etc. It also describes many of the properties of liquids, and of the continuous passage of a gas to the liquid condition. The constants a and b must be evaluated from observations. The numerical values of the "constants" in van der Waals' equation are not always quite constant at different temperatures. Quite a number of attempts have been made to still further modify the gas equation so that it may describe the behaviour of gases under wide variations of pressure and temperature. J. D. van der Waals (1888) found that for carbon dioxide, $R = 0.00369$; $b = 0.0023$; and $a = 0.00874$. Let us find how van der Waals' equation describes the behaviour of carbon dioxide under variations of pressure. This gas is known to deviate

¹ *E.g.* suppose AB represents two units, and the diameter of the molecule $\frac{1}{10}$ unit; the molecule then moves through 1.9 units of space; and in AC through 0.90 unit, not 0.95 unit, in passing from one boundary wall to the other.

² Certain assumptions introduced into the kinetic theory have led to the view that b is relatively near four times the material volume of the molecule; the assumptions have not been so firmly established that there is no room for doubt, and some consider that b represents the real volume of the molecule; others that b is much greater than 4 times the volume of the material nucleus, and that b is not constant but varies somewhat with changes of temperature and pressure. A molecule seems to be a complex vibratory system with a material nucleus which requires b volumes of space in which to perform its oscillatory movements. Hence, b is sometimes called the "vibratory volume" of the molecule.

considerably from the $pv = 1$ required by Boyle's law. On comparing the values of pv for carbon dioxide, calculated from the equation, at 20° :

$$\left(p + \frac{0.00874}{v^2}\right)(v - 0.0023) = 1.08$$

with the numbers observed by E. H. Amagat (1893) at 20° , we get:

p atmospheres.	pv .	
	Observed.	Calculated.
1	1.000	1.000
50	0.680	0.678
75	0.180	0.179
100	0.228	0.226
200	0.219	0.411
500	0.938	0.936

The agreement between theory (van der Waals' equation) and fact (the observed data) is quite good. It will be remembered that if the gas behaves according to the equation $pv = RT$, pv would have the same value for all pressures. As a matter of fact, the value of pv first decreases and then increases for all gases except hydrogen and helium. The two corrections act in opposite ways. At first the value of pv is decreased by the molecular attraction, and increased by the finite dimensions of the molecule. At low pressures, the correction for molecular attraction preponderates over that required for the volume of the molecule; while the correction for the volume of the molecules is relatively large when the volume of the gas is compressed very small by a large pressure. (Cf. pp. 99, 105.) Several attempts have been made to still further improve the gas equation by the introduction of other terms, involving special constants which have to be evaluated from the experimental numbers. They are therefore of very limited application.

§ 16. The Critical State of Gases.

The ordinary gaseous and liquid states are only widely separated forms of the same condition of matter, and may be made to pass into one another by a series of gradations so gentle that the passage shall nowhere present any interruption or breach of continuity. Gas and liquid are only distant stages of a long series of continuous physical changes.—T. ANDREWS.

In 1869 T. Andrews found that if carbon dioxide be gradually compressed in a vessel suitable for the observation, the volume diminishes more rapidly than would occur if Boyle's law correctly described the behaviour of the gas; and when the pressure attains a certain value, the gas begins to liquefy. A further decrease in the volume does not change the pressure, but only increases the quantity of gas liquefied. At length, when all the gas has liquefied, a large increase of pressure only causes a minute decrease in the volume of the liquid, since liquids in general undergo but a small change of volume on compression.

If the experiment be made with carbon dioxide at 0° , the gas commences to liquefy when the pressure has attained 35.4 atmospheres; if at 13.1° , liquefaction commences at 48.9 atmospheres pressure; if at 30° , at 70 atmospheres pressure; while if the temperature exceeds 31° , no pressure, however great, will liquefy the gas. Other gases exhibit similar phenomena. For each gas there is a particular temperature above which liquefaction is impossible, however great be the applied pressure. Andrews called this the critical temperature of the gas. For instance, the critical temperature of:

Hydrogen.	-241°	Nitrous oxide	$+35^\circ$
Nitrogen	-146°	Ammonia	$+130^\circ$
Oxygen	-119°	Sulphur dioxide	$+155^\circ$
Carbon dioxide	$+31^\circ$	Water	$+368^\circ$

The pressure required to liquefy the gas at the critical temperature is called the critical pressure; and the volume occupied by unit volume of gas at n.t.p. when the critical temperature and pressure are attained is

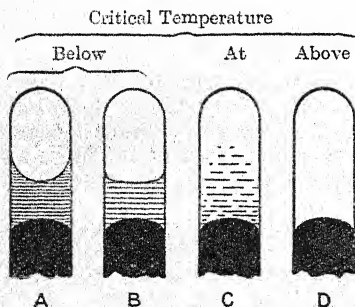


FIG. 261.—Diagrammatic Illustration of the Critical State of a Gas.

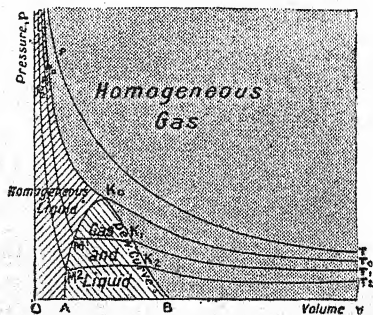


FIG. 262.— $p : v$ —Curves for Carbon Dioxide.

called the **critical volume**. It is interesting to notice the influence of temperature on carbon dioxide, partly liquid, partly gaseous. Fig. 261, A, represents the upper end of a glass tube in which the partly liquefied carbon dioxide is confined over mercury, at 18° . The surface of the liquid has a sharply defined curved meniscus. On raising the temperature, the meniscus of the liquid becomes flatter and flatter, Fig. 261, B, until, at 31° , Fig. 261, C, the surface seems to disappear. The sharp line of demarcation between the liquid and the gas vanishes. At 40° , the tube contains a homogeneous gas, Fig. 261, D. Liquid carbon dioxide cannot exist at this temperature, however great the pressure.

The relation between the pressure and the volume of, say, carbon dioxide, at different temperatures— T , T_0 , T_1 , T_2 —is represented diagrammatically in Fig. 262. The portion of the curve K_2T_2 , or K_1T_1 , represents the behaviour of the gas when no liquid is present; the portion K_2M_2 , or K_1M_1 , the behaviour of the gas in the presence of its own liquid; and M_2p_2 , or M_1p_1 , the behaviour of the liquid when no gas is present. It will be observed that K_2M_2 or K_1M_1 is a straight line horizontal with the v -axis. It illustrates in a graphic manner the well-known law: At

any fixed temperature, the pressure of a gas in the presence of its own liquid is always the same. The curve $T_0K_0p_0$ represents the relation between pressure and volume at the critical temperature; and the curve T , the relation between p and v at a temperature when the gas does not liquefy. The line $K_0K_1K_2B$ represents the condition under which the gas, compressed at the stated temperatures T_0 , T_1 , T_2 , begins to liquefy, and it is hence called the dew curve, because a gas under a gradually increasing pressure first shows signs of liquefaction under conditions represented by a point on this line; similarly, the line $K_0M_1M_2A$ is called the boiling curve, because a liquid, under a gradually diminishing pressure, first shows signs of vaporization under conditions represented by a point on this line. Note also that the lines K_0A , K_0B , and K_0p_0 , divide the plane of the paper into three regions. Every point to the right of BK_0p_0 represents a homogeneous gas; every point in the region AK_0B represents a heterogeneous mixture of gas and liquid; and every point to the left of AK_0p_0 , a homogeneous liquid. The gas in the region K_0BVT_0 is below its critical temperature, and is then said to be a vapour—*vide* p. 187. The diagram, Fig. 262, thus represents the conditions of equilibrium of a liquid or a gas under different conditions of temperature, pressure or volume.

It is interesting to note historically that Caignard de la Tour, long before Andrews' experiment, noticed that when a liquid is heated in a sealed tube there is a definite temperature at which the surface of separation between the gas and liquid disappeared, and the whole contents of the tube become homogeneous. Caignard de la Tour's experiments thus demonstrate that the critical temperature is the upper limit to the liquid state; and Andrews' experiments prove that the critical temperature is the lower limit to the gaseous state. The passage from the one state to the other proceeds in a continuous manner. The liquid and gaseous states are continuous, not abrupt. The properties—density, surface tension, viscosity, refractive power, heat of vaporization, compressibility, etc.—of a liquid gradually lose their distinctive character as the temperature is raised, until, at the critical temperature, the properties of liquid and gas are the same. There is no evidence of a change in molecular structure when, say, carbon dioxide passes from the one state of aggregation to the other; nor is there any evidence of a polymerization of the molecules¹ when the common gases condense to liquids. The difference between liquids and gases *below* the critical temperature seems to be a question of molecular attraction.

§ 17. Carbon Monoxide—Preparation and Occurrence.

Molecular weight, CO = 28. Melting point, -203° ; boiling point, -190° ; critical temperature, -139.5° . Relative vapour density ($H_2 = 2$), 27.81; (air = 1), 0.9672.

Two oxides of carbon are very well known—carbon monoxide and carbon dioxide. In the section dealing with producer gas, p. 873, it will be shown that the former is partially burnt or oxidized carbon, and is

¹ Nitrogen peroxide, water, and some other substances appear to polymerize and form compound molecules on passing from the gaseous to the liquid state of aggregation. The properties of the condensing gases do not then exhibit that continuity shown by carbon dioxide and gases which do not polymerize.

formed when carbon is heated in a limited supply of air: $2C + O_2 = 2CO$. It is also produced when carbon dioxide is heated in the presence of an excess of carbon: $CO_2 + C = 2CO$. Carbon dioxide is reduced in a similar manner when passed over hot metallic zinc: $CO_2 + Zn = ZnO + CO$. If zinc oxide be reduced with carbon, metallic zinc and carbon monoxide are produced: $ZnO + C = Zn + CO$; so also with litharge, lead monoxide, PbO , we have $PbO + C = Pb + CO$. F. de Lassone first made carbon monoxide by this reaction in 1776, and Priestley prepared it by heating iron oxide with charcoal in 1796. Lavoisier knew that this gas burned to carbon dioxide, but he was not able to satisfactorily fit carbon monoxide in with his theory of oxidation. This became possible when W. Cuickshank, in 1800, showed that the gas was nothing but a "gaseous oxide of carbon," and F. Clément and J. B. Désormes, in 1801, made clear the composition of the gas.

Carbon monoxide seldom occurs free in nature, and then only in small quantities. Minute traces have been detected in air, in volcanic gases, occluded in coal, and in meteorites. It is found in tobacco smoke, in chimney gases where the air is not in a sufficiently large excess, and in the gases from blast and other furnaces.

Preparation.—Carbon monoxide is best made in the laboratory by heating concentrated sulphuric acid to about 100° in a flask (Fig. 263), fitted with a double-bored rubber stopper; one hole of the stopper is fitted with a gas delivery tube, and the other with a tap funnel containing concentrated (98 per cent.) formic acid, $H.COOH$. The gas can be washed by bubbling it through a concentrated solution of sodium hydroxide. The tip of the funnel dips below the surface of the liquid. Each drop of formic acid produces some bubbles of gas. The reaction is simple: $H.COOH + H_2SO_4 = H_2SO_4.H_2O + CO$. There is no frothing, and the velocity of the stream of gas can be regulated by the rate at which formic acid is dropped on to the sulphuric acid. Concentrated sulphuric acid can also be mixed with sodium formate in a flask provided with a safety funnel, and very gently warmed. Carbon monoxide is evolved. The rate of evolution of the gas depends on the concentration of the acid.

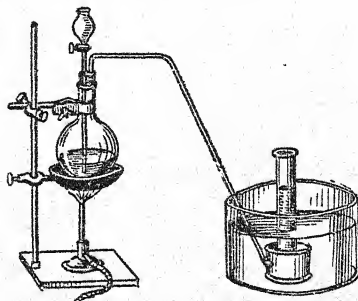


FIG. 263.—Preparation of Carbon Monoxide.

The gas is also made by gently heating 20 to 30 grams of crystalline oxalic acid, $(COOH)_2.2H_2O$, just covered with concentrated sulphuric acid. The rate of evolution of the gas is determined by the temperature of the mixture. Equal volumes of carbon dioxide and carbon monoxide are produced: $(COOH)_2.2H_2O + H_2SO_4 = H_2SO_4 + 3H_2O + CO + CO_2$. The carbon dioxide is removed by passing the gas through a series of wash-bottles containing a solution of caustic soda. This process is more troublesome than the formic acid process. The same remark applies to the preparation of carbon monoxide by heating, say, 30 grams of crystalline potassium ferrocyanide (or cyanide) with 300 c.c. of concentrated sulphuric acid in a capacious flask. In this case, the temperature of decomposition is high, and if the mixture be heated too quickly, the gas is apt to come off so

quickly as to be beyond control. The reaction is symbolized: $K_4FeCy_6 + 6H_2SO_4 + 6H_2O = 2K_2SO_4 + 3(NH_4)_2SO_4 + FeSO_4 + 6CO$. If dilute sulphuric acid be used, hydrocyanic acid as well as carbon monoxide may be formed: The reaction in the latter case is represented: $2K_4FeCy_6 + 3H_2SO_4 \rightarrow 3K_2SO_4 + K_2Fe_2Cy_6 + 6HCy$.

§ 18. The Properties of Carbon Monoxide.

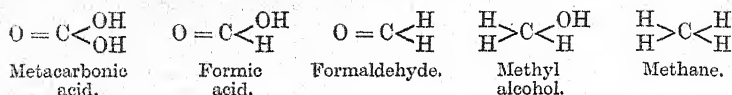
Carbon monoxide is a colourless, tasteless, and odourless (poisonous) gas. One volume of water dissolves about 0.03 volume at 0° , and about 0.023 volume at 20° . The gas condenses to a colourless transparent liquid at -190° , at atmospheric pressure; the liquid solidifies at -203° . When a lighted taper is plunged into the gas, the taper is extinguished (non-supporter of combustion), but the gas is inflamed (combustible), and burns with a blue flame. The blue lambent flame which appears on the surface of a clear red coke (or coal) fire, or over the ignited lime in a lime kiln, is carbon monoxide. If a beaker be held over a burning jet of carbon monoxide for a moment, and clear lime-water be poured into the beaker, a turbidity shows the presence of carbon dioxide. Ozone can oxidize carbon monoxide to carbon dioxide. Carbon monoxide, on burning, forms carbon dioxide: $2CO + O_2 = 2CO_2$. Collect a mixture of two volumes of carbon monoxide and one volume of oxygen in a soda-water bottle; apply a lighted taper to the mouth of the bottle; the mixture explodes with some violence. If the gases be thoroughly dried by means of phosphorus pentoxide a spark can pass through the mixture of two volumes of carbon monoxide and one volume of oxygen without explosion. There appears to be a slight combination, but only in the path of the spark. H. B. Dixon showed that the presence of gases containing hydrogen—hydrogen sulphide, pentane, ammonia, etc.—act like water, and make the mixture explosive; whereas gases not containing hydrogen—sulphur dioxide, carbon disulphide, etc.—are inert, for the mixture does not explode. Water thus appears to act as a catalytic agent, and Dixon thinks that there is a cyclic process (p. 164) involving the reactions:



This emphasizes how even the so-called simple reactions are much more interesting than the ordinary chemical equations would lead us to suppose. How much more wonderful must be the reactions which are now represented by more complex equations!

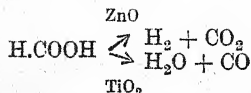
The reduction and hydration of the carbon oxides.—If moist carbon monoxide be passed over palladium black in the cold, formic acid, $H.CO.OH$, can be detected among the products of the action. The end products of the reaction are hydrogen and carbon dioxide—the hydrogen is largely retained by the palladium. If the flame of burning carbon monoxide be allowed to play upon ice, formic acid can be detected in the water formed by the melting ice. Hence it is assumed that the first stage of the reaction is $H_2O + CO = H.CO.OH$; followed by $H.CO.OH = H_2 + CO_2$. The formic acid is an intermediate product of the two consecutive reactions (p. 165). The reverse reaction, reduction of carbon dioxide to formic acid: $CO_2 + H_2 \rightleftharpoons H.CO.OH$, occurs under the influence of palladium black. It might also be added that formic acid and formaldehyde have been produced by exposing moist carbon dioxide to the ultra-violet radiations from a mercury

lamp. The relationship between formic acid, metacarbonic acid, and formaldehyde will appear from the graphic formulæ :

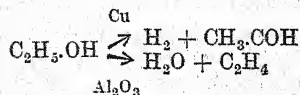


Carbon monoxide is reduced by hydrogen to methane, CH_4 , if the mixture be passed over finely divided nickel at about 200° : $\text{CO} + 3\text{H}_2 = \text{H}_2\text{O} + \text{CH}_4$.

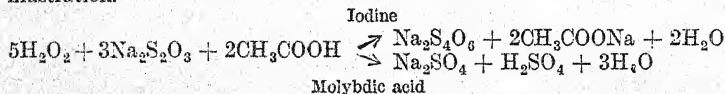
A catalyst can often direct a chemical transformation along one path in preference to another (J. W. Mellor, 1904), and this view is amply supported by P. Sabatier's work on the catalysis of organic compounds (1914). Totally different effects may be produced by catalysts with almost identical physical properties. For instance, the vapour of formic acid is decomposed wholly into hydrogen and carbon dioxide when passed over zinc oxide, but into water and carbon monoxide when passed over titanous oxide. Symbolically



Again, ethyl alcohol, $\text{C}_2\text{H}_5\text{OH}$, is converted into aldehyde, CH_3COH , and hydrogen by finely divided copper, and into ethylene and water by alumina.



The different effects of hydrogen peroxide upon sodium thiosulphate according as iodine or molybdic acid are employed as catalysts, is a further illustration.



Carbon monoxide reduces ammoniacal cuprous sulphate solution and ammoniacal silver nitrate; it has no action on rosaniline decolorized by sulphurous acid. Carbon monoxide is not absorbed by solutions of potassium or sodium hydroxide (distinction from carbon dioxide). Unlike carbon dioxide, carbon monoxide is soluble in a saturated solution of cuprous chloride in concentrated hydrochloric acid, or of cuprous chloride in ammonia. Hence these solutions are used as absorbents for carbon monoxide in gas analysis. Some suppose that the carbon monoxide unites with the cuprous chloride, forming a compound: $2\text{CuCl} \cdot \text{CO} \cdot 2\text{H}_2\text{O}$.

Action on iron.—Carbon monoxide is an important reducing agent: ferric oxide, for instance, is reduced to the metal in the blast furnace: $\text{Fe}_2\text{O}_3 + 3\text{CO} = 2\text{Fe} + 3\text{CO}_2$. The action is, however, somewhat complex

and dependent upon the temperature. When metallic iron is heated in a current of carbon monoxide at about 330° , the carbon monoxide is decomposed, forming ferrous oxide and carbon: $\text{Fe} + \text{CO} = \text{FeO} + \text{C}$ (+ 20.1 Cals.). At higher temperatures the ferrous oxide is reduced to the metal by this gas: $\text{FeO} + \text{CO} = \text{Fe} + \text{CO}_2$. At 300° the former reaction is about twelve times as fast as the second. As the temperature is raised, the speed of the latter reaction is accelerated until it appears as if the ferric oxide were directly reduced by the carbon monoxide; any ferrous oxide formed is at once reduced. This is a curious example of consecutive reactions discussed on p. 163. T. Graham showed that metallic iron absorbs about four times its volume of carbon monoxide at a dull red heat, and the gas is given off when the metal is heated in air. Hence iron appears to be permeable to carbon monoxide, and if this gas be passing up a hot stove pipe, with the flue gases, some carbon monoxide could pass through the metallic iron into the atmosphere of the room.

Composition.—A mixture of carbon monoxide and oxygen can be exploded in the apparatus illustrated in Fig. 209. It will be found that 100 vols. of carbon monoxide with 100 vols. of oxygen gave 150 vols. of gas. When the gaseous product of the explosion was treated with a solution of potassium hydroxide, 50 vols. of oxygen remained. The absorption was conducted as indicated on p. 809, Fig. 253. This shows that 100 vols. of carbon monoxide combined with 50 vols. of oxygen to form 100 vols. of carbon dioxide. By Avogadro's hypothesis, this agrees with the equation: $2\text{CO} + \text{O}_2 = 2\text{CO}_2$. The density of the gas confirms this, for it is nearly fourteen times that of hydrogen. Hence the molecular weight of the gas is 28 (hydrogen = 2). The atomic weight of carbon is 12, and oxygen 16, and $12 + 16 = 28$. Hence CO , not C_2O_2 , C_3O_3 . . . is the correct formula for carbon monoxide. Alternative graphic formulæ were discussed on page 88. The study of the refractive power of substances containing bivalent oxygen led J. W. Brühl to attribute the exceptional behaviour of carbon monoxide to its containing quadrivalent oxygen.

Carbon monoxide—a poison.—Carbon monoxide is an active poison. When it is respired in the lungs, it unites with the hæmoglobin of the blood, forming a bright cherry-red coloured compound carbonyl-hæmoglobin which prevents the hæmoglobin performing its regular work of oxidizing waste tissue (p. 906). Some consider it acts by paralyzing the nerve centres which govern respiration. Less than 1 per cent. in the atmosphere is sufficient to cause death when inhaled. J. S. Haldane says that air containing $\frac{1}{50}$ of one per cent. of carbon monoxide will produce giddiness on exertion if breathed for from one-half to two hours; and $\frac{1}{10}$ per cent. makes one unable to walk; $\frac{1}{5}$ per cent. leads to loss of consciousness and perhaps death; $\frac{1}{2}$ per cent. means probable death; and 1 per cent. will lead to unconsciousness in a few minutes, followed before long by death.¹ A. Moso found that animals could live in oxygen, containing 6 per cent. of carbon monoxide, at 2 atm. pressure. Fatal accidents have arisen from the breathing of carbon monoxide, when its presence has not

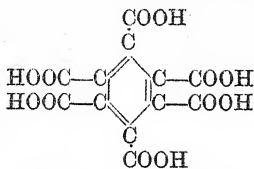
¹ First aid treatment for carbon monoxide poisoning, or "gassing," as it is often called, is artificial respiration at once, accompanied by the use of oxygen for about 10 minutes. A person seriously gassed should be kept warm and protected from cold air; he should not exert himself by walking; and if there is a tendency to fainting, a little brandy or whisky should be administered.

been suspected. It is formed when a gas flame is exposed to a cold surface as might occur in some "patent water heaters" used in ill-ventilated bathrooms, etc.; in slow combustion stoves, and charcoal pans. It is often present in rooms where gas producers are at work (p. 874); and in rooms with an escape of coal gas, water gas, etc. The poisonous character of "after-damp" in coal mines is due to the presence of some carbon monoxide along with the carbon dioxide formed during an explosion of fire-damp or coal dust; in air after the use of explosives for blasting; underground fires; etc. J. S. Haldane recommends miners to carry a mouse in an open cage in a suspected atmosphere. As soon as the mouse shows signs of sluggishness or exhaustion, a dangerous amount of carbon monoxide is probably present.

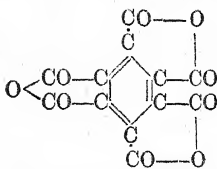
Testing for carbon monoxide.—The detection of carbon monoxide is not easy. (1) A piece of white filter paper moistened with *platinum* or *palladium chloride*— PtCl_4 or PdCl_2 —is turned pink, green, or black, owing to the reduction of the chloride to the metal. The intensity of the tint is proportional to the amount of carbon monoxide in the given atmosphere (freed from ammonia and hydrogen sulphide). The tint of test-papers saturated with the metal chloride solutions and exposed to the given atmosphere for a definite period, can be compared with papers standardized with atmospheres containing known amounts of gas, and an estimate of the carbon monoxide made accordingly. Some organic substances produce the same reaction, and mistakes have been accordingly made. (2) Iodine pentoxide is reduced by carbon monoxide: $5\text{CO} + \text{I}_2\text{O}_5 = \text{I}_2 + 5\text{CO}_2$. This reaction is a sensitive test for the gas. The liberated iodine gives a violet solution with chloroform or carbon disulphide. (3) *Vogel's blood test* is considered the most reliable. If ordinary blood be diluted with 200 times its volume of water, a yellowish-red solution is obtained, with blood containing carbon monoxide, the solution is distinctly pink. J. S. Haldane recommends placing a mouse in the suspected atmosphere for, say, fifteen minutes. It is then killed. A known volume of the blood drawn from the heart of the mouse is then systematically diluted, simultaneously with another sample of ordinary oxygenated blood, until the tints of both solutions appear the same in daylight. A comparison of the amount of blood in the two solutions will give a rough idea of the amount of carbon monoxide in the air, since it is found that the blood of the mouse 20 per cent. saturated represents 0.02 per cent. of carbon monoxide; 33 per cent. saturated with 0.04 per cent. carbon monoxide; 50 per cent. saturated with 0.08 per cent. of carbon monoxide; 66 per cent. saturated with 0.16 per cent. carbon monoxide in the air.

Carbon suboxide.—Certain other oxides of carbon have been reported by B. C. Brodie (1873), M. Berthelot (1876), etc.: (1) Brodie's oxide, C_4O_3 , is formed as a reddish-brown mass by exposing carbon monoxide to the prolonged action of the electric discharge; (2) Meyer and Steiner's oxide (1913), C_4O_3 , was obtained by heating mellitic acid, $\text{C}_6(\text{COOH})_6$, with benzoyl chloride, $\text{C}_6\text{H}_5\text{COCl}$, in a flask fitted with a reflux condenser for six hours, and found that colourless crystals of a substance with the same empirical formula as Brodie's oxide, but with the probable molecular formula, C_{12}O_9 . It appears to be formed by the abstraction of three molecules of water from mellitic acid: $\text{C}_6(\text{COOH})_6 = 3\text{H}_2\text{O} + \text{C}_{12}\text{O}_9$. The oxide is almost insoluble in cold water, but is rehydrated to mellitic

acid by hot water—hence its name mellitic anhydride. The usual graphic formulæ for the acid and the anhydride are :



Mellitic acid, $\text{C}_6(\text{COOH})_6$



Mellitic anhydride, C_{12}O_9

(3) Berthelot's oxide, C_3O_2 , is formed as a dark brown mass by heating Brodie's oxide between 300° and 400° ; and (4) Diels' carbon suboxide, C_3O_2 , has been made by O. Diels (1906) by distilling malonic acid, $\text{CH}_2(\text{COOH})_2$, or its (ethereal) salts with a large excess of phosphorus pentoxide under reduced pressure (12 mm. mercury) at about 300° . The gases are cooled so as to remove the unchanged acid, carbon dioxide, ethylene, etc. The reaction is represented: $\text{CH}_2(\text{COOH})_2 = 2\text{H}_2\text{O} + \text{C}_3\text{O}_2$. When passed through a tube dipping in liquid air, a white solid with a pungent odour is obtained. The solid melts between -107° and -108° and boils at 7° . When treated with water, carbon suboxide re-forms malonic acid, and hence this oxide is regarded as malonic anhydride. The suboxide is combustible and burns with a blue smoky flame, forming carbon dioxide. The vapour and liquid polymerize at ordinary temperatures, forming a dark red solid soluble in water. The suboxide decomposes rapidly at 37° and instantaneously at 100° . Analysis by combustion with copper monoxide, by explosion with oxygen, and vapour density determinations by Hofmann's method all correspond with the formula C_3O_2 . The constitution is uncertain, it is generally supposed to be represented by $\text{O}=\text{C}=\text{C}=\text{O}$, that is, carbon dicarbonyl, $\text{C}(\text{CO})_2$, or by $\text{O}=\text{C} \begin{array}{c} \diagup \text{C} \diagdown \\ \text{O} \end{array} \text{O}$. An oxide, C_3O_2 ,

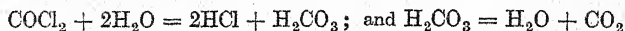
has also been formed by the action of the silent electric discharge on carbon monoxide: $4\text{CO} = \text{C}_3\text{O}_2 + \text{CO}_2$. There is little doubt about the existence of carbon suboxide C_3O_2 , but the identity of the other oxides— C_5O_3 and C_4O_3 —is not well established.

§ 19. Carbonyls.

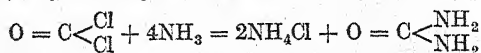
According to the valency hypothesis, oxygen a dyad and carbon a tetrad, carbon monoxide is an unsaturated compound $=\text{C}=\text{O}$ with two sleeping valencies, p. 89. This is in agreement with the ready combination of carbon monoxide with oxygen to form carbon dioxide $\text{O}=\text{C}=\text{O}$. When a mixture of carbon monoxide and sulphur vapour is passed through a moderately hot tube, a gaseous carbonyl sulphide— COS —analogous with carbon dioxide is formed and hence also called carbon oxysulphide. In carbonyl sulphide one atom of sulphur has taken the place of one atom of oxygen in carbon dioxide so that its formula is written: $\text{S}=\text{C}=\text{O}$. When CO acts as a radicle, as it appears to do in many reactions, it is called carbonyl. Carbon dioxide can thus be regarded as a carbonyl oxide.

Carbonyl chloride.—Carbon monoxide also directly unites with chlorine in sunlight to form *phosgene*—in allusion to its formation under the influence of light—or carbonyl chloride COCl_2 . The same gas is produced when a

mixture of chlorine and carbon monoxide is passed over bone charcoal (catalytic agent). The gas liquefies at 8°. The formation of the gas is conveniently shown by filling two equal-sized gas cylinders, by displacement, one with chlorine, and one with carbon monoxide. Place the vessels mouth to mouth and allow the gases to mix thoroughly. Both cylinders show the characteristic colour of chlorine. Cover each cylinder with a glass plate. Let one be exposed to a strong light for a short time, the colour of the chlorine rapidly disappears, and when the glass plate is removed, the contents of the cylinder fume strongly owing to the decomposition of the carbonyl chloride in contact with the moisture of the air. The gas also decomposes by hydrolysis in contact with water, forming hydrochloric acid and carbon dioxide:



When heated with metal sulphides to about 400°, carbonyl sulphide and a metal chloride is formed: $\text{COCl}_2 + \text{MS} = \text{COS} + \text{MCl}_2$. With an excess of ammonia (NH_3), carbonyl chloride forms urea— $\text{CO}(\text{NH}_2)_2$. Thus:



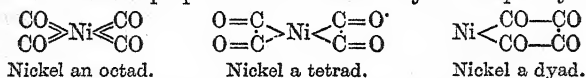
The ammonium chloride— NH_4Cl —is insoluble in alcohol, urea is soluble, hence the two compounds are easily separated.

Metabolic products.—Urea is a waste product of animal life. Here then, by purely chemical processes, we can build up from the proper elements a compound which is formed by living animals. Wöhler's synthesis of urea, by another process, in 1826, appears to have attracted a lot of attention because the teachers of chemistry—Gmelin, Berzelius, etc.—then taught that compounds formed by animals and plants were produced by the action of a vital force; that "within the sphere of living nature, the elements obey laws totally different from those which obtain in inanimate nature," and that "organic substances cannot be prepared artificially." Wöhler's synthesis rendered it necessary to reconsider these assertions. Wöhler communicated his discovery to Berzelius in these words: "I must tell you that I can prepare urea without requiring a kidney or an animal, a man or a dog." Hundreds of different animal and vegetable products have since been prepared in the laboratory, and the hypothesis that a peculiar vital force is necessary for the preparation of these products has been abandoned. The growth of the organized structure of animals and plants must not be confused with the formation of chemical compounds.

About the middle of the seventeenth century chemistry was subdivided into "organic" and "inorganic." The latter term was intended to denote the chemistry of the inorganic substances found in the "mineral kingdom," as contrasted with the chemistry of the compounds produced as the result of vital processes in animal and vegetable life. The sharp distinction between organic and inorganic compounds lost its clearness when chemists succeeded in bulking up typical "organic" compounds from "mineral elements." The term "organic chemistry" still persists, but its original significance is gone. To-day, the term is really understood to mean the chemistry of *certain* carbon compounds. Some carbon compounds like iron carbide (steels), etc., are always treated in inorganic chemistry; other compounds—e.g. calcium carbide—are arbitrarily included or excluded from organic chemistry according to the writer's whim or caprice.

Nickel tetracarbonyl.—L. Mond and C. Langer (1890) discovered that

when carbon monoxide is passed over finely divided metallic nickel¹ at between 40° and 50°, and the gas then passed through a cold tube, a colourless liquid condenses. This boils at 43° under a pressure of 751 mm. of mercury; and at -25° solidifies to a mass of needle-shaped crystals. The gas decomposes at 180° into metallic nickel and carbon monoxide. The gas burns with a luminous flame, and a deposit of black metallic nickel, resembling soot, is obtained when the flame is brought in contact with a cold porcelain dish. If the gas be passed through a hot glass tube, dust, or a mirror-like deposit of metallic nickel is formed on the glass. Mond utilizes this reaction for extracting nickel from its ores. Two analyses of the compound furnished 34.26 and 34.33 per cent. of nickel. These numbers correspond with the formula, $[\text{Ni}(\text{CO})_4]_n$. The vapour density of the gas, air = 1, is 6.01; and if hydrogen = 2, the density of air is 28.75; consequently, the vapour density of nickel carbonyl (hydrogen = 2) is $6.01 \times 28.75 = 172.94$. The vapour density corresponding with $\text{Ni}(\text{CO})_4$, by Avogadro's hypothesis, is 170.68. Hence the formula of nickel carbonyl is written $\text{Ni}(\text{CO})_4$. Some consider that the "CO" in nickel carbonyl is a dyad, and the nickel an octad, because of certain physical properties,² and write the formula as indicated below; others consider both the nickel and the CO to be dyads. Several other formulæ have been suggested—*e.g.* nickel may be regarded as a tetrad. The cyclic formula is in harmony with the general behaviour of these compounds, for it shows how the properties of the metal may be completely masked.



As the student acquires familiarity with graphic chemical formulæ, he will find that the arrangement of a number of atoms into a graphic formula, consistent with the valencies of the elements, is sometimes rather a test of man's ingenuity than a representation of the actual grouping of the atoms. In some cases, too, it is thought necessary to assume special valencies for particular combinations. When it is not possible to decide between several conflicting formulæ, select the most probable, and use the simplest—hypotheses must not be multiplied without necessity. This important rule is sometimes called Occam's razor.

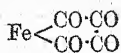
Iron carbonyl.—Cobalt does not form a carbonyl under the same conditions as nickel.³ If iron be treated like nickel, at 80°, a gas—iron tetracarbonyl, $\text{Fe}(\text{CO})_4$ —is formed, which burns with a yellow flame. The reaction with iron is very much slower than with nickel. At ordinary temperatures the finely divided iron furnishes iron pentacarbonyl— $\text{Fe}(\text{CO})_5$ —which condenses to an amber-yellow liquid boiling at 103° and freezing at -21°. Its vapour density (air = 1) is 6.5. Unlike nickel carbonyl, this compound is decomposed by light, forming gold-coloured crystals of iron enneacarbonyl— $\text{Fe}_9(\text{CO})_6$. This latter compound decomposes, when

¹ The nickel is best made by reducing nickel oxalate in a current of dry hydrogen at about 300°.

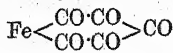
² Principally the refractive index, which is from three to four times as great as is usually exhibited by dyad nickel salts.

³ This of course does not mean that a cobalt carbonyl cannot be prepared, because $\text{Co}_2(\text{CO})_8$ can be made if the reaction occurs under great pressure—150 atmospheres—in the form of orange-yellow crystals. Colourless crystals of molybdenum carbonyl— $\text{Mo}(\text{CO})_6$ —are known. Several other carbonyls are known.

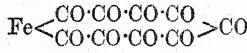
heated, forming carbon monoxide, and iron pentacarbonyl. The composition of the three carbonyls may be represented :



Tetracarbonyl.



Pentacarbonyl.



Enneacarbonyl.

Lamp glasses used for gas burners sometimes appear to redden owing to the deposition of a very thin film of iron oxide on the inner surface ; and accordingly it is inferred that a trace of a volatile iron carbonyl is present in the illuminating gas. Iron carbonyl is also supposed to be formed in iron cylinders in which "water gas" (p. 877) has been stored under pressure for some time.

§ 20. Carbon Sulphides.

Molecular weight, $\text{CS}_2 = 76.14$. Melting point, -110° ; boiling point, 46° . Specific gravity of liquid at 0° , 1.292. Vapour density ($\text{H}_2 = 2$), 77.06; (air = 1), 2.68.

Carbon disulphide was discovered by W. A. Lampadius while studying the action of pyrites on carbon. Clement and Desormes rediscovered the

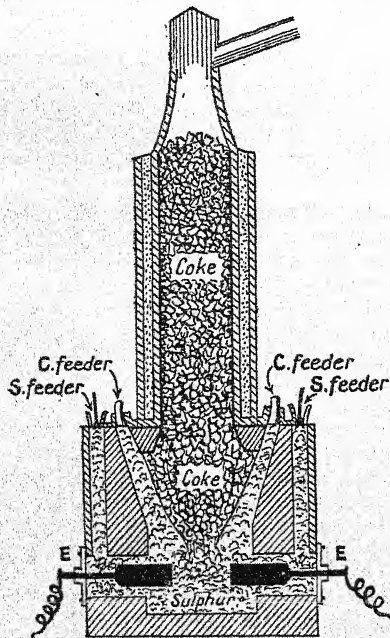


FIG. 264.—Taylor's Carbon Disulphide Furnace.

liquid in 1802. When sulphur vapour is passed over red-hot charcoal, the two elements combine, forming volatile carbon disulphide: $\text{C} + \text{S}_2 = \text{CS}_2 - 28.7$ Cals. This reaction is therefore endothermal, and it bears some analogies with the exothermal reaction $\text{C} + \text{O}_2 = \text{CO}_2 + 96.98$ Cals. The volatile compound of carbon and sulphur is condensed in vessels surrounded with cold water. The product is contaminated with free sulphur, which volatilizes with the carbon disulphide; some hydrogen sulphide is formed at the same time by the action of sulphur on the hydrogen in the charcoal.

Manufacture of carbon disulphide.—In the manufacturing process, the charcoal is heated in vertical cast-iron or earthenware retorts set in a suitable furnace. The heat of the furnace also melts the sulphur placed in a vessel near the base of the retort, the sulphur vapour rises through the red-hot charcoal and forms carbon disulphide which

escapes at the top. The carbon disulphide is condensed in long condensing coils—30 feet long.

In Taylor's electric process (1899) a cylindrical furnace 40 feet high and

16 feet in diameter is packed with coke from the top, Fig. 264; the coke is renewed through the side hopper *C*. An alternating current is sent through the electrodes *E* set at right angles to one another at the base of the furnace. The heat melts the sulphur on the floor of the furnace; the vapour of sulphur rises through the coke, forming carbon disulphide. Fresh sulphur is introduced through the hopper as shown in the diagram. The carbon disulphide passes off at the top of the furnace, and is condensed in the condensing coils. The electrical process is practically continuous and is free from troublesome leakages and heat losses incidental to the retort process. The crude product by either process can be purified by repeated redistillation and a prolonged agitation with mercury and porous calcium chloride, followed by decantation and a slow distillation in the dark; but the manufacturing process of purification is a "trade secret."

Properties.—Carbon disulphide, if pure, is a colourless, mobile, refractive liquid, with an aromatic smell not at all displeasing, although the smell of commercial carbon disulphide is usually disagreeable and rancid. Light slowly decomposes the pure liquid in such a way that a small quantity of some product—possibly a polymer—is formed which raises the boiling point of the liquid and produces an unpleasant smell. Air also seems to affect it, particularly if the temperature is raised a little. So sensitive is purified carbon disulphide to these influences that it has been said: pure carbon disulphide is an ephemeral product. If breathed constantly, in small quantities, carbon disulphide is injurious to health, and in large quantities, fatal. Its specific gravity at 0° is 1.292. It freezes at -116° ; melts at -0.86° ; and boils at 46.27° . 100 grams of water at 0° dissolve 0.258 gram of carbon disulphide; and at 20° , 0.101 gram. Most of the alkali salts are insoluble in carbon disulphide; several salts of the metals are sparingly soluble in this liquid—thus, mercuric nitrate and halides are slightly soluble since, at 18° , 100 parts of a saturated solution contain 0.03 gram of mercuric chloride; 0.12 gram of mercuric bromide; 0.22 gram of mercuric iodide; and at 25° these numbers become respectively 0.05, 0.23, and 0.38. Carbon disulphide mixes in all proportions with alcohol, ether, benzene, and essential oils. It is also a good solvent for sulphur, phosphorus, iodine, bromine, camphor, gums, resins, waxes, fats, and caoutchouc; and it is largely employed in the industries on account of its solvent properties. It is also used as an insecticide.

Thiocarbonic acid.—A compound called thiocarbonic acid, H_2CS_3 , analogous with carbonic acid, H_2CO_3 , is known. Thiocarbonic acid is a reddish-yellow oil formed by the action of dilute hydrochloric acid on thiocarbonates. It is very unstable, but some of its salts—thiocarbonates—are fairly stable. They are formed by the action of carbon disulphide on the sulphides—e.g. calcium thiocarbonate: $\text{CS}_2 + \text{CaS} = \text{CaCS}_3$; or on the hydroxides: $3\text{Ca}(\text{OH})_2 + 3\text{CS}_2 = 2\text{CaCS}_3 + \text{CaCO}_3 + 3\text{H}_2\text{O}$. Hence CS_2 may be regarded as the anhydride of thiocarbonic acid in the same sense that CO_2 is the anhydride of carbonic acid. Potassium thiocarbonate is used as an insecticide for vine-lice.

Thiocarbonyl chloride, CSCl_2 .—If carbon disulphide and chlorine be left in contact for some weeks, particularly if a little iodine be present, a compound, CSCl_2 , is formed, but the best way of making this substance is to heat a mixture of phosphorus pentachloride and carbon disulphide in

a sealed tube at 100° . A reaction symbolized: $\text{PCl}_5 + \text{CS}_2 = \text{PSCl}_3 + \text{CSCl}_2$, seems to occur. The compound thiocarbonyl chloride, CSCl_2 , is related with carbonyl chloride, COCl_2 . Thiocarbonyl chloride is a foetid-smelling liquid which boils with a little decomposition at 149° . The liquid is slightly hydrolyzed by water, and it polymerizes to $(\text{CSCl}_2)_2$ as shown by its effect on the freezing point of benzene. Normal thiocarbonyl chloride is probably constituted: $\text{S} = \text{C} = \text{Cl}_2$, and the polymeride $\text{S} = \text{CCl} - \text{S} - \text{C} \equiv \text{Cl}_2$.

Carbon monosulphide, CS .—The analogy between the behaviour of sulphur and oxygen in many chemical compounds makes it appear highly probable that a carbon monosulphide, CS , will exist and will bear the same relation to carbon monoxide, CO , that carbon disulphide bears to carbon dioxide. The isolation of carbon monosulphide has proved unexpectedly difficult. The reaction between nickel carbonyl and thiocarbonyl chloride furnishes a brown solid which is either carbon monosulphide, CS , or a polymer, C_nS_n . The reaction is represented: $n\text{CSCl}_2 + n\text{Ni}(\text{CO})_4 = 4n\text{CO} + n\text{NiCl}_2 + \text{C}_n\text{S}_n$. The brown product is soluble in concentrated sulphuric acid, and it is precipitated unchanged on dilution. Carbon monosulphide is not produced when carbon disulphide acts on nickel carbonyl. When electric sparks are passed through the vapour of carbon disulphide at a low pressure, a mixture of carbon monosulphide and unchanged carbon disulphide is obtained. The former polymerizes to the brown solid as the temperature rises. A bright red liquid carbon subsulphide, C_2S_2 , has been made by passing the vapour of carbon disulphide through a quartz tube at 1000° to 1100° ; and by exposing the vapour of carbon disulphide to the electric arc. The chief interest in this compound turns on its possible analogy with carbon suboxide, C_3O_2 . If an arc-discharge between electrodes of a mixture of tellurium and graphite, beneath carbon disulphide, the same subsulphide is formed, and a volatile carbon sulphotelluride, CSTe . Both compounds are unstable. The latter melts at -54° . With selenium in place of tellurium, carbon sulphoselenide, SCeS , is formed as a yellow liquid boiling at 84° .

Questions.

1. How many tons of salt and of sulphuric acid containing 70 per cent. of real acid are required to make 200 tons of salt cake? Assuming that the manufacturer condenses only 90 per cent. of the hydrochloric acid evolved, what weight of this gas is allowed to escape?—*New Zealand Univ.*

2. 2.4 litres of air at 0°C . and 76 cm. pressure are shaken with 50 c.c. of $\frac{1}{100}$ normal $\text{Ba}(\text{OH})_2$ solution. After removal of the precipitate, the remaining alkali requires 35 c.c. of $\frac{1}{100}$ normal acid to neutralize it. Find the volume of carbonic anhydride in 10,000 volumes of the air.—*St. Andrews Univ.*

3. Give the chief chemical and physical properties of CO_2 and CS_2 .—*Owens Coll.*

4. Chlorine, sodium carbonate, and caustic soda are obtained from common salt. Describe fully the chemical reactions involved in the production of these compounds.—*Univ. North Wales.*

5. If hydrogen chloride and carbonic anhydride are found mixed with air and it is desired to separate, one at a time, the substances in the mixture until but one remains, what would be the successive methods employed?—*Amherst Coll., U.S.A.*

6. The absorption coefficient of nitrogen dissolved in water is 0.0152 at 12.6° . What volume of the gas measured at standard temperature and pressure is absorbed by one litre of water at 12.6° , at each of the following pressures: 1000 mm., 748.2 mm., 391 mm., and 14.3 mm. ?—*New Zealand Univ.*

7. Explain how the qualitative and quantitative composition of carbon dioxide and carbon monoxide have been ascertained.—*St. Andrews Univ.*

8. You are given a mixture of calcium carbonate and calcium sulphate. How would you proceed to determine the proportion of each present?—*Aberystwyth Univ.*

9. At 20° and 760 mm., a volume of carbonic anhydride is measured and found to occupy 75 litres. The gas is passed over red-hot carbon and after being brought to the same conditions of temperature and pressure as above, is again measured. What volume is the gas found to occupy?—*Amherst Coll., U.S.A.*

10. Apply the law of chemical mass action to the liberation of carbon dioxide from a carbonate by an acid.—*Sheffield Scientific School, U.S.A.*

11. How can it be shown that the gas obtained by heating marble is identical with that evolved when marble is acted upon by dilute hydrochloric acid? Calculate the volume of carbon dioxide measured at 15° and 740 mm. pressure obtainable from 10 grams of calcium carbonate ($\text{Ca} = 40, \text{O} = 16, \text{C} = 12$, one litre of hydrogen at N. T. P. weighs 0.09 gram).—*Sheffield Univ.*

12. Give a full account of what you learned about sulphurous anhydride by experiment. Contrast sulphites with carbonates, and sulphurous anhydride with carbonic anhydride.—*Princeton Univ., U.S.A.*

13. What is meant by the law of mass action? State what you understand by the term "active mass." Discuss from the standpoint of the mass action law, (a) the dissociation of calcium carbonate by heat, (b) the effect produced by adding hydrochloric acid to an aqueous solution of sodium chloride.—*St. Andrews Univ.*

14. The following sentence appears in a modern text-book of chemistry: "Men striving to bring individual chemical processes to the highest state of perfection by utilizing all waste products." Illustrate this statement by two examples of the recovery of valuable chemical substances from waste products, and explain in detail the chemical reactions involved.—*London Univ.*

15. Enunciate Gay-Lussac's law of gaseous volumes. What change in volume would take place in one litre of oxygen by (a) burning sulphur in it, (b) converting 6 per cent. of it into ozone, (c) combining it with twice its volume of carbon monoxide? Accurately describe the way in which you would carry out any one of these experiments.—*London Univ.*

16. What distinction can be drawn between dissociation and decomposition? Illustrate your answer by reference to calcium carbonate, and explain why the presence of free carbon dioxide diminishes the extent to which the dissociation of the substance can proceed at a given temperature.—*London Univ.*

17. The air of a room was tested for carbon dioxide by drawing 100 litres of it through weighed bulbs, containing caustic potash. The temperature was 15° and the pressure 750 mm. The increase in weight of the bulb was 0.08 gram. What was the percentage by volume of carbon dioxide in the air of the room?—*London Univ.*

18. Explain and illustrate the meaning of the expression *coefficient of solubility* of a gas. What is meant by the statement that ammonia at ordinary temperatures deviates from Henry's law of solubility?—*London Univ.*

19. The acid formed by dissolution of carbon dioxide in water might have its structure represented by the formula $\text{CO}(\text{OH})_2$ or $\text{C}(\text{OH})_4$. State fully (a) the grounds on which the existence of carbonic acid in such a solution may be assumed, and (b) the evidence in favour of both these formulæ.—*Board of Educ.*

20. How are carbonyl, phosphoryl, and sulphuryl chlorides respectively prepared? Describe briefly their properties, and show how their characteristic reactions help to elucidate the constitution of the corresponding acids.—*Board of Educ.*

21. If 12 grams of carbon disulphide were burned in 100 litres of air at 14° and 740 mm., what would be the volume of the gaseous mixture at N. T. P. (a) when the combustion was complete, and (b) when the mixture had been extracted by concentrated caustic potash? How could the relative proportions of the two dioxides be determined? (Air contains 21 per cent. of oxygen, and 1 litre of hydrogen at N. T. P. weighs 0.09 gram.)

22. How do you account for the fact that in spite of the large amount of oxygen consumed in respiration and combustion, the percentage of oxygen in the atmosphere remains practically unaltered?—*Victoria Univ., Manchester.*

23. What would you observe and what chemical changes (if any) would take place if the following elements were heated strongly in (a) a current of air, (b) a current of steam:—copper, mercury, arsenic, zinc, charcoal?—*London Univ.*

24. On passing pure carbon monoxide over copper oxide, it was found that the loss of weight was 24.36 grams and that the amount of carbon dioxide formed was 67.003 grams. From these data calculate the atomic weight of carbon ($O = 16$). How would you make the carbon monoxide required for this experiment, and ascertain its freedom from impurity?—*Board of Educ.*

25. What are the common impurities in ordinary water, and to which of them is hardness due? Outline methods by which hardness can be (a) reduced, (b) removed, and explain why curd is formed when soap is used with hard water.—*Sheffield Univ.*

26. Describe the principal features of the "mechanical" system of water filtration. What is the filter substance in this process, and how is it produced? Give equation.—*Cornell Univ., U.S.A.*

27. How are temporarily hard waters formed in nature? How are they softened? Account for the formation of caves in limestone rock. How are the stalactites and stalagmites formed in these caves? What is formed when carbon dioxide gas is passed into a solution of potassium hydroxide? Write all equations.—*Princeton Univ., U.S.A.*

28. Stas found that when the lower oxide of carbon was passed over heated copper oxide, 63.64 grams of carbonic anhydride were formed and that the copper oxide lost in weight 22.515 grams. Stas had found previously that 1.232 grams of carbon when burned in oxygen yielded 4.517 grams of carbonic anhydride. Calculate the percentage composition of (1) carbonic anhydride, (2) the lower oxide of carbon; and show that the result is in agreement with the law of multiple proportions.—*Punjab Univ.*

29. A liquid formed by fermentation is found to consist of carbon, hydrogen, and oxygen. On analysis, 0.375 gram of the liquid yielded 0.715 gram of carbon dioxide and 0.441 gram of water. Calculate the percentage composition and probable formulae of the liquid.—*New Zealand Univ.*

30. The spring water of our chalk districts is much harder than the river water of the same localities. Why is this the case? At Canterbury the water was rendered soft by the addition of a certain quantity of lime-water. Explain the process.—*London Univ.*

31. It may be stated that with the same amount of oxygen, carbon monoxide contains twice as much carbon as carbon dioxide. State any facts which may be supposed to support this assertion, and without quoting the commonly accepted formulae of these gases, discuss the question by reference to their known properties.—*Science and Art Dept.*

32. Explain what occurs when carbon monoxide reacts with (a) ammoniacal cuprous chloride, (b) chlorine, (c) nickel, (d) iron. Describe the properties of the several products and the conditions under which they are obtained. Have any of these products received technical applications, and for what purposes, if any, is each thus applied?—*Board of Educ.*

33. What is meant by the critical point in connection with the compression of gases? Give a diagram showing the general form of the isothermals of a substance, such as water or carbon dioxide, at and near the critical temperature, and considerably above and below that temperature.—*Sydney Univ.*

34. Explain the following results: 25 c.c. of $\frac{1}{10}N - Na_2CO_3$ required 25 c.c. of $\frac{1}{10}N - HCl$ to neutralize it using methyl orange as the indicator, but 12.5 c.c. using phenolphthalein as indicator; but 25 c.c. of the $\frac{1}{10}N - HCl$ required 25.2 c.c. of the $\frac{1}{10}N - Na_2CO_3$ solution to neutralize it using methyl orange and 48.4 c.c. using phenolphthalein.—*Sydney Univ.*

35. How can you estimate accurately the amount of carbon dioxide in atmospheric air? What causes tend (1) to augment, and (2) to diminish the proportion in air?—*Science and Art Dept.*

36. Water at 15° can absorb 756 times its volume of ammonia. What mass of ammonia can be dissolved in one litre at 15° under a pressure of 10 atmospheres, the volume of one gram of ammonia at 0° and a pressure of 760 mm. being 1317 c.c.?—*Sydney Univ.*

CHAPTER XXXVI

HYDROCARBONS

§ 1. Methane—Occurrence, Preparation, and Properties.

Molecular weight, $\text{CH}_4 = 16.03$. Melting point, $= -184^\circ$; boiling point, -160° . Relative vapour density ($\text{H}_2 = 2$), 15.95; (air = 1), 0.5457.

Occurrence.—In 1667 T. Shirley published an account of his observations on the combustible gas which dribbled from a ditch near Wigan (Lancashire). The district was underlain with coal. The formation of carburetted hydrogen in the decomposition of organic matter was afterwards reported by many others, and methane has been recognized as a distinct individual since 1805, when W. Henry proved that the gas was

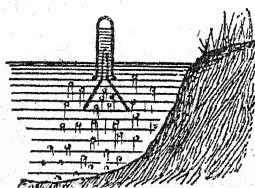


FIG. 265.—Marsh Gas.

different from the other hydrocarbons known at that time. In 1776 A. Volta distinguished methane from inflammable air in that the former alone produced carbon dioxide when burnt. C. L. Berthollet (1785) showed that the gas contained carbon and hydrogen. The inflammable gaseous emanations from mines, etc., were mentioned by Pliny at the beginning of our era. Methane is often called marsh gas because it is sometimes generated in stagnant marshy pools. The bubbles of gas which rise to the surface, when the mud at the bottom of a pond is disturbed, often contain methane. The gas can be collected in many stagnant ponds by the use of an inverted test-tube full of water and fitted with a funnel as indicated in Fig. 265. The funnel directs the bubbles, disturbed by poking a stick into the mud at the bottom of the pond, into the test-tube.¹ The test-tube must be securely clamped or it may overturn when full of gas. Methane appears to be a product of the gradual decay of vegetable matter in a very limited supply of air. There are many unreliable descriptions by inaccurate observers (p. 6) of the luminous phenomena sometimes seen in marshy grounds, and called *ignes fatui*, or *will-o'-the-wisps* or *jack-o'-lanterns*—the *feux follets* of the French, and the *Irrlichter* of the Germans—are generally supposed to be produced by the ignition of marsh gas derived from the decay of vegetable matter. The decomposition of animal matters may result in the escape of some phosphuretted hydrogen which

¹ The writer, when a boy, had an attack of typhoid fever supposed to have been received as a result of this experiment in a miasmatic swamp.

oxidizes when diluted with air, producing a luminous glow, probably the source of the so-called *corpse candles* and the *death-light*.

Methane occurs absorbed or occluded by coal sometimes under considerable pressure. When the pressure is relieved—*e.g.* when a face of coal is exposed in mining, or during a sudden fall in atmospheric pressure—this gas escapes from coal, sometimes in a continuous stream with a hissing sound—called by the miner a “singer” or a “blower.” Methane, mixed with more or less air and carbon dioxide, is common in the atmosphere of coal mines, and hence this gas is found in the air discharged by the “upcast” ventilating shaft. On account of its inflammable nature, the mixture of gases occluded in coal is called *firedamp*. Other synonymous terms are “light carburetted hydrogen gas” or simply “gas.” So far as it is possible to tell by analysis, firedamp (air absent) contains between 80 and 98 per cent. of methane, up to 4 per cent. of carbon dioxide, and up to 11 per cent. of nitrogen. Marsh gas and firedamp are therefore more or less impure forms of methane. Methane is formed in considerable quantities when coal is heated out of contact with air. Coal gas may contain 30 to 40 per cent. of methane. Enormous quantities of gas, containing 80 to 98 per cent. of methane, escape from the petroleum springs in Baku and the Caucasus (Russia), and in the oilfields of Indiana, Ohio, Pennsylvania, etc., where it is called *natural gas*. The percentage composition (volume) of typical samples from Ohio and from Baku is reported as follows:

	Hydrogen. ¹	Methane.	Ethane.	Carbon dioxide.	Nitrogen.	Oxygen.
Baku (Russia) .	0.9	93.1	3.3	2.2	0.5	0
Ohio (U.S.A.) .	1.9	92.8	0.4	0.8	3.8	0.3

Preparation.—Hydrogen slowly unites with carbon at a dull red-heat to form methane. Mayer and Altmayer (1907) found that the higher the temperature the smaller the proportion of methane; thus, at 500°, the equilibrium mixture contains 37.47 per cent. of hydrogen and 62.53 per cent. of methane; and at 800°, 95.5 per cent. of hydrogen and 4.4 per cent. of methane. The reaction is so slow in the absence of catalytic agents that there is some doubt if equilibrium was attained. The catalytic action of nickel is explained on the assumption that nickel hydride is formed as an intermediate product; and the catalytic action of nickel oxide is explained by assuming that it is reduced to nickel which then forms the intermediate hydride. W. Spatoeff (1913) gives the alternate hypothesis that the oxide is reduced by the carbon, forming carbon dioxide as an intermediate product: $2\text{NiO} + \text{C} = \text{CO}_2 + 2\text{Ni}$; the hydrogen reduces the carbon dioxide to methane and water: $\text{CO}_2 + 4\text{H}_2 = \text{CH}_4 + 2\text{H}_2\text{O}$; and the water is then reduced by the metal to form nickel oxide and hydrogen: $\text{Ni} + \text{H}_2\text{O} = \text{H}_2 + \text{NiO}$. Hence, a small quantity of nickel oxide can produce a large quantity of methane. In confirmation, a mixture of carbon dioxide and hydrogen at about 420° in the presence of nickel gave relatively large amounts of methane. Thus a mixture of hydrogen (82 per cent.) with carbon dioxide (6.5 per cent.) at 13 atm. pressure gave a mixture containing 74.2 per cent. of methane and 8 per cent. of hydrogen.

Methane is usually prepared by the following process: Ten grams of

¹ Some claim that free hydrogen does not occur in natural gas, and that the alleged presence of hydrogen in these gases is due to a mistake in the analysis. That question must be left with the experts in natural gas.

fused sodium acetate are intimately mixed with three times that weight of soda lime in a mortar. Introduce the dry mixture in a glass or copper flask, or in a copper tube retort. Close the copper flask *A*, Fig. 266, with a cork to which a delivery tube is attached. The flask is strongly heated, and when all the air is expelled, the methane is collected over water in the usual manner. The reaction in the retort or flask is usually represented by the equation: $\text{CH}_3\text{COONa} + \text{NaOH} \rightarrow \text{Na}_2\text{CO}_3 + \text{CH}_4$. Soda lime or barium oxide is used in place of sodium hydroxide because of the fusibility of the latter. Soda lime is a mixture of calcium hydroxide and sodium hydroxide.

The gas prepared by this process is not very pure, but the mode of preparation is useful for ordinary purposes. For instance, this gas burns with a luminous flame, whereas pure methane burns with a non-luminous flame.

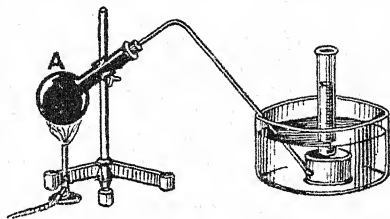


FIG. 266.—Preparation of Methane.

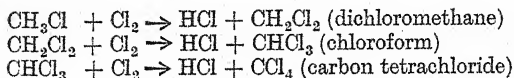
An impure gas—mixed with ammonia and acetylene—is formed by the action of water on commercial aluminium carbide: $\text{Al}_4\text{C}_3 + 12\text{H}_2\text{O} \rightarrow 4\text{Al}(\text{OH})_3 + 3\text{CH}_4$. The gas can be washed free from acetylene by means of an ammoniacal solution of cuprous chloride. When pure methane is required, methyl iodide, CH_3I , is reduced by nascent hydrogen formed by the action of a mercury: aluminium couple, or a copper: zinc couple on water or alcohol, thus: $\text{CH}_3\text{I} + 2\text{H} = \text{CH}_4 + \text{HI}$. This method of preparation is discussed in works on organic chemistry.

Properties.—Pure methane is free from colour, and is without smell. The gas prepared by the ordinary process usually has a slight smell, but this is due to the presence of impurities. Methane is lighter than air. 100 volumes of water at 0° dissolve $5\frac{1}{2}$ volumes of the gas; and at 20° , $3\frac{1}{2}$ volumes. Methane is rather more soluble in alcohol. It liquefies at 0° under a pressure of 140 atmospheres. The liquid boils at -160° , and solidifies at -184° . Methane has no well-defined physiological action on the system other than diluting the oxygen and so inducing suffocation.

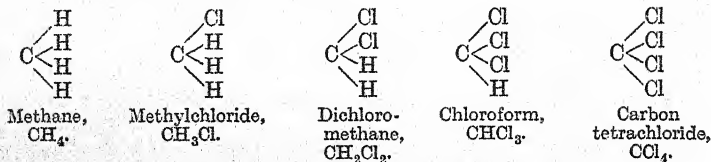
Combustion and explosion of methane.—When a lighted taper is plunged into a cylinder of this gas, held mouth downwards, the taper is extinguished (non-supporter of combustion), and the gas burns at the mouth of the jar (combustible). The gas, if pure, burns with a pale blue non-luminous flame, forming carbon dioxide and water: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$. The gas ignites in air at a temperature between 650° and 750° . If the supply of air be limited, some hydrogen and carbon monoxide will be formed. One volume of methane requires two volumes of oxygen, *i.e.* 9.5 volumes of air for complete combustion. Such a mixture is violently explosive. A red-hot wire will ignite the moist mixture. The explosive effect diminishes in violence with increasing proportions of air—a mixture of one volume of methane with 19 volumes of air is explosive; and in the presence of coal dust, W. Galloway says that one volume of methane mixed with 111 volumes of air is explosive. It must be added that coal dust itself is explosive when mixed with air. If a mixture of hydrogen, methane, and air, with an excess of oxygen, be led over palladium-asbestos,

the hydrogen alone burns—fractional combustion—the methane is not acted upon provided the temperature of the palladium does not rise above 100°. There is no explosion. Hence follows a method for estimating the amount of hydrogen in a mixture of hydrogen and methane. For the detection of methane, see "Flame caps," p. 913.

Action of chlorine.—When a mixture of equal volumes of chlorine and methane is exposed to *diffused daylight*, they gradually react, forming methyl chloride— CH_3Cl . Thus: $\text{CH}_4 + \text{Cl}_2 \rightarrow \text{HCl} + \text{CH}_3\text{Cl}$. If more chlorine be present, the chlorine gradually replaces all the hydrogen, forming:



This process of replacing one or more atoms in a molecule by equivalent atoms is called substitution. We have:



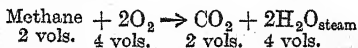
By the action of concentrated sodium hypoiodite on alkaline solutions of acetone, carbon tetraiodide is formed, and carbon tetrabromide or tetrachloride can be produced in a similar way. By heating chloroform with bromine in sealed tubes a series of chlorobromides— CCl_3Br , CBr_2Cl_2 , CBr_3Cl , CBr_4 —have been obtained. If an excess of chlorine reacts with methane in *direct sunlight* an explosion occurs with the separation of carbon: $\text{CH}_4 + 2\text{Cl}_2 \rightarrow 4\text{HCl} + \text{C}$. Shake a mixture of one-third volume of methane with two-thirds of a volume of chlorine in a gas cylinder; apply a lighted taper. Acid fumes of hydrogen chloride will be formed, and soot will be deposited in the cylinder.

All the hydrogen can thus be expelled from methane in four stages, and the carbon only in one stage, hence it is probable that the molecule of methane contains one atom of carbon and four atoms of hydrogen. Similarly, in dealing with, say, hydrogen chloride: (1) only one compound is known; (2) the molecules of hydrogen and chlorine are halved during the formation of two molecules of hydrogen chloride; and (3) the hydrogen and chlorine can only be expelled from hydrogen chloride each in one stage. Hence it is inferred that the molecule of hydrogen chloride contains one atom of chlorine and one atom of hydrogen; and that each of the molecules of chlorine and hydrogen contain two atoms. This purely chemical evidence agrees with the formulae based on Avogadro's hypothesis. This argument was used in deducing the basicity of acids, p. 345.

Carbon tetrachloride, CCl_4 , is rather important commercially as a solvent for fats—degreasing woollen and linen fabrics, bones, oleaginous seeds, etc. It can replace benzene, which is dangerously inflammable. Carbon tetrachloride is a colourless non-inflammable liquid of specific gravity 1.63, boiling at 77°. The liquid attacks metals particularly if moisture

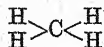
be present owing to the formation of hydrochloric acid. It is made by passing dry chlorine through carbon disulphide, CS_2 , containing a little iodine (as catalyst) in suspension—asbestos impregnated with magnesium chloride is also used in place of iodine— $\text{CS}_2 + 3\text{Cl}_2 = \text{S}_2\text{Cl}_2 + \text{CCl}_4$. On distillation the tetrachloride can be removed from the sulphur chloride. It is purified by washing it in dilute sodium hydroxide and redistillation from bleaching powder.

Composition.—If a measured volume of methane be mixed with an excess of air or oxygen, and exploded in a eudiometer, the contraction in volume determines the amount of hydrogen present, since the corresponding amount of water condenses to a liquid whose volume is negligibly small in comparison with the gas. The carbon dioxide can be absorbed by potassium hydroxide and the corresponding contraction represents the amount of carbon dioxide formed. In illustration, 10 c.c. of methane were mixed with 40 c.c. of oxygen in a Hempel's burette, Fig. 209. The mixed gases were driven into the Hempel's explosion pipette, and exploded. The gases were returned to the burette, and the volume measured 30 c.c. instead of the original 50 c.c. The contraction was therefore 20 c.c. Water vapour contains its own volume of hydrogen, hence 10 volumes of methane contains the equivalent of 20 volumes of hydrogen. Again, the gases were transferred from the burette to the absorption pipette charged with potassium hydroxide solution, Fig. 253. On returning the gases to the burette, the volume measured 20 c.c. The contraction due to the absorption of the carbon dioxide was 10 c.c., and the 20 c.c. excess oxygen remained in the burette. One volume of carbon dioxide is equivalent to one volume of oxygen and one atom of carbon. Hence the analysis has furnished the following data :



showing that the atoms of carbon and hydrogen must be in the ratio 1 : 4 ; or the formula is $(\text{CH}_4)_n$. The relative density of methane (air = 1) is 0.559 ; and for hydrogen = 2, we have $28.755 \times 0.559 = 16.07$. If $n = 1$, the vapour density of CH_4 will be $12 + 4 = 16$. Hence the formula for methane is CH_4 .

If hydrogen be univalent, the only possible plane graphic formula is :

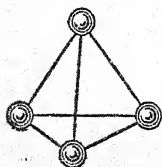


which makes carbon quadrivalent. The usual "proof" that all four hydrogen atoms are arranged in equivalent positions in the molecule proceeds by first preparing compounds like CH_3Cy , $\text{CH}_2\text{Cy}(\text{COOH})$, $\text{CHCy}(\text{COOH})_2$, and subsequently replacing the acid groups by hydrogen. Thus CH_3Cy gives CHHHCy ; $\text{CH}_2\text{Cy}(\text{COOH})$ gives CHHCyH ; and $\text{CHCy}(\text{COOH})_2$ gives CHCyHH . All these products, CH_3Cy , are the same, and hence it is assumed that at least three atoms of hydrogen in methane are equivalent, or else there is only one stable form for compounds of the type CH_3Cy . There are several other analogous "proofs." Organic chemists can give a number of reasons for assuming that the carbon atom behaves as if it had the form of an equilateral tetrahedron, each apex representing a free valency. Thus, if the four valencies of the carbon

atom be limited to one plane, and chlorine be univalent, it ought to be possible to make the two isomers (*cf.* p. 796):



In spite of numerous attempts this has not been accomplished. So far as we can tell, all four valencies of carbon are equivalent and symmetrical, or else one of the above compounds is very unstable and, when formed, immediately passes into the stable modification. The former hypothesis is much the more probable. *The four valencies can be symmetrical and equivalent only if they are directed from a central carbon atom towards the four corners or faces of a regular tetrahedron.* The carbon atom thus behaves as if it were shaped like a tetrahedron, and nothing is more surprising, said H. E. Armstrong, than the completeness with which the vast array of facts included in organic chemistry may be ordered by reference to the tetrahedron model. A graphic representation of the real or imaginary molecule of methane is shown in the adjoining sketch. The spheres represent atoms of hydrogen. Some consider that the formula of methyl chloride, CH_3Cl , would be better represented by a tervalent chlorine, $\text{H}_2=\text{C}=\text{Cl}-\text{H}$. This is mentioned to show that dogmatic statements cannot yet be made. The tetrahedron of modern organic chemistry recalls the views of the ancient Greek philosophers for the poly-



Methane— CH_4 .

hedron of Plato (*Timæos*), and the mystic quaternion of Pythagoras played prominent parts in their views on matter.

Attempts to isolate the hydrocarbon methylene, CH_2 , containing bivalent carbon have proved futile; the corresponding carbon monoxide may have a bivalent carbon or quadrivalent oxygen and hence the evidence is ambiguous. A study of the constitution and properties of the alkyl isocyanides, $\text{R}-\text{N}=\text{C}$; of hydrocyanic acid, $\text{NH}=\text{C}$, and the cyanides, $\text{M}-\text{N}=\text{C}$; of fulminic acid, $\text{HO}-\text{N}=\text{C}$, and the fulminates, $\text{MO}-\text{N}=\text{C}$; and of the monohalogen derivatives of acetylidene, *e.g.* $\text{IH}=\text{C}=\text{C}$, and the dihalogen derivatives, *e.g.* $\text{I}_2=\text{C}=\text{C}$, have led organic chemists to favour the hypothesis that these compounds contain bivalent carbon. This latter hypothesis is not improbable when it is remembered that other members of the oxygen family of elements—*viz.* sulphur, selenium, and tellurium—exist as bi- and sexi-valent atoms.

§ 2. Ethylene—Occurrence, Preparation, and Properties.

Molecular weight, $\text{C}_2\text{H}_4 = 28.03$. Melting point, -169° ; boiling point, -103° . Relative vapour density, $28.12 (\text{H}_2 = 2)$ and $0.978 (\text{air} = 1)$.

Ethylene was discovered by J. J. Becher about 1669, and its properties were distinguished from ordinary inflammable air by the Dutch chemists J. P. Deimann, A. P. van Troostwijk, N. Bondt, and A. Lauwerenburgh (1795). These chemists also showed that the gas contained carbon and hydrogen. It was clearly distinguished from methane in 1805 by W. Henry.

Occurrence.—Ethylene occurs in natural gas (p. 848). Some analyses of "pit gases" show that up to 6 per cent. of this gas may occur in the air

in coal pits. This gas is also obtained when coal or wood is heated in closed vessels and coal gas contains from 4 to 10 per cent. of ethylene.

Preparation.—The gas is most conveniently prepared by the action of dehydrating agents (sulphuric acid, zinc chloride, or phosphoric acid) upon alcohol— C_2H_5OH . For this purpose put 50 c.c. of syrupy phosphoric acid in a 250 c.c. flask furnished with a two-hole rubber stopper fitted with a tap-funnel drawn out into a capillary end, and also a thermometer, T (Fig. 267). The side neck of the flask is connected with a delivery tube leading to a wash-bottle, and finally to the gas trough. The flask is heated to about 200° on a sand-bath or metal plate. Ethyl alcohol is slowly run from the tap-funnel below the surface of the phosphoric acid. The alcohol is decomposed into water and ethylene. The water is retained by the phosphoric acid. The reaction is represented in symbols: $C_2H_5OH \rightarrow H_2O + C_2H_4$. The reaction with sulphuric acid is singularly interesting. At ordinary temperatures, alcohol, C_2H_5OH , unites with sulphuric acid to form ethylsulphuric acid, $C_2H_5(HSO_4)$, thus, if the liquid be kept cold, $C_2H_5OH + H_2SO_4 \rightleftharpoons H_2O + C_2H_5.HSO_4$ represents the state of the system. At about 60° , the sulphuric acid acts as a catalytic agent, and by a cyclic series of reactions involving the temporary formation of ethyl sulphuric acid as an intermediate compound, which immediately unites with more alcohol to form ether, $(C_2H_5)_2O$, boiling at 34.6° : $C_2H_5OH + C_2H_5.HSO_4 = H_2SO_4 + (C_2H_5)_2O$, and regenerating the sulphuric acid. In this reaction—sometimes called *Williamson's reaction*—the formation of ether, the intermediate compound can be isolated. If the mixture of sulphuric acid and ethyl alcohol, or rather of ethyl sulphuric acid, be rapidly heated to about 150° , it breaks down into ethylene and sulphuric acid: $C_2H_5.HSO_4 = H_2SO_4 + C_2H_4$.

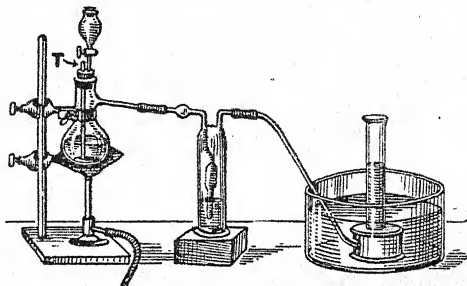


FIG. 267.—Preparation of Ethylene.

Another method of preparation is to pass the vapour of alcohol over alumina heated to about 360° —above this temperature ethylene is decomposed, ether may be formed, and the catalyst alumina becomes clogged with tarry matters which cannot be readily burned off without destroying the efficiency of the catalyst; thoria is not so readily contaminated as alumina, and the tarry matters can be burned off without interfering with the activity of the catalyst. Other methods of preparation are described in text-books on organic chemistry.

Properties.—Ethylene is a colourless gas with a peculiar ethereal odour. It has nearly the same density as air. 100 volumes of water at 0° dissolve 25.7 volumes of the gas, and at 20° , 15 volumes; the gas is nearly 13 times as soluble in alcohol. Ethylene liquefies at 0° under a pressure of 43 atmospheres; the liquid boils at -103° , and solidifies at -169° . The gas behaves like methane towards a lighted taper, but it burns with a luminous smoky flame unless it be diluted with hydrogen or methane.

Ethylene is decomposed at a high temperature. This is illustrated by passing the gas through a bulb-tube of hard glass. When all the air has been expelled, heat the bulb in the blowpipe flame. By rotating the bulb, a mirror-like deposit of carbon can be formed in the interior of the bulb—Fig. 268. One volume of ethene requires three volumes of oxygen, or 14.3 volumes of air for complete combustion. Such a mixture is a powerful explosive. The explosion is more violent than with methane and air.

If the tube contains aluminium turnings, heated not quite to the melting point of the metal, the hydrocarbon is decomposed into its elements, hydrogen is evolved, and carbon remains partly deposited mechanically on the surface of the metal, and partly united with the metal as carbide. If a mixture of hydrogen and ethylene be passed over reduced copper, between 180° and 300° , the ethylene, C_2H_4 , is reduced to ethane, C_2H_6 , a gas resembling methane in many of its properties. Finely divided nickel can also be used as a catalyst. The reaction commences (slowly) at about 30° . P. Sabatier argued that the nickel acts by forming "an unstable hydride such as NiH_2 , which fixes the hydrogen temporarily and transfers it to the hydrogenizable substances." In the presence of compounds capable of

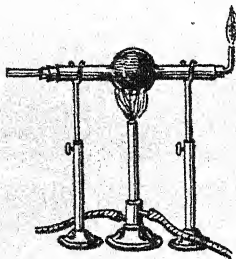
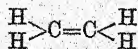


FIG. 268.—Decomposition of Ethylene.

giving up hydrogen, the nickel can take it from them, and to eliminate it in the gaseous state, and by altering the conditions, nickel can be made to promote dehydrogenation as well as hydrogenation. Thus, aldehyde, $CH_3CO.H$, and an excess of hydrogen, passed over nickel at 200° , forms alcohol: $CH_3CH_2.OH$; and conversely, alcohol passed over nickel at 250° forms aldehyde and hydrogen: $CH_3CO.H + H_2 \rightleftharpoons CH_3CH_2.OH$. The reaction with hydrogen and ethylene is also reversible: $C_2H_4 + H_2 \rightleftharpoons C_2H_6$. The dehydrogenation of ethane, C_2H_6 , is best observed by passing the gas over blue tungstic oxide, between 300° and 350° , with nickel at this temperature, the ethylene deposits carbon and forms methane with some hydrocarbons which condense to liquids at ordinary temperatures.

The composition of the gas can be determined by volumetric analysis as in the case of methane. The result shows that ethylene is $(C_2H_4)_n$. The vapour density ($H_2 = 2$) is nearly 28. Hence the formula for the gas must be C_2H_4 . The graphic formula for ethylene with carbon quadrivalent and hydrogen univalent is not possible if all the valencies have to be saturated or "satisfied." Hence the graphic formula involves two sleeping or unsaturated valencies. On joining the carbon atoms by a double bond, we get:



One molecule of ethylene readily combines with chlorine, forming an oily liquid which is ethylene dichloride— $C_2H_4Cl_2$. Half fill a tall cylinder with ethylene, carefully but quickly fill up the cylinder with chlorine, and allow the cylinder to remain in the trough; the water gradually rises in the cylinder, the ethylene dichloride will be seen floating on the surface

of the water. Hence ethylene was once called *olefiant gas*. The oily ethylene dichloride was once termed *Dutch liquid*. Similarly with bromine, a litre flask filled with ethylene and 2 c.c. of bromine will form a colourless oil of $C_2H_4Br_2$, ethylene dibromide, on the bottom of the flask. Fill a cylinder one-third with ethylene, and two-thirds with chlorine. Mix the gases well, and apply a light; acid fumes are formed with much soot: $C_2H_4 + 2Cl_2 = 4HCl + 2C$. The ethylene molecule also combines directly with sulphuric acid, hydrogen bromide, etc. The compounds of ethylene with chlorine, bromine, etc., are called **addition products**.

In gas analysis, ethylene is absorbed in a Hempel's pipette charged with fuming sulphuric acid, or with bromine, and these reagents can be employed to remove ethylene—and *unsaturated hydrocarbons* generally—from a mixture of ethylene, with hydrogen, methane, air, oxygen, and other gases not absorbed by this reagent. Methane is not absorbed unless it is left standing in contact with the fuming sulphuric acid for a long time.

§ 3. Acetylene—Preparation and Properties.

Molecular weight, $C_2H_2 = 26.02$. Melting point, -81.5° ; boiling point, -83.6° . Relative vapour density, 26.46 ($H_2 = 2$); 0.92 (air = 1).

This gas was discovered by E. Davy in 1836. He obtained it by the action of water on the black mass formed as a by-product in the manufacture of potassium.

Preparation.—M. Berthelot prepared it by sparking carbon electrodes in a current of hydrogen gas in a glass bulb as indicated in the diagram—Fig. 269. The carbon and hydrogen unite directly: $2C + H_2 = C_2H_2$. Pure hydrogen unites directly with carbon at temperatures exceeding 1100° . At 1200° , about 0.35 per cent. of methane is formed; at 1500° , 0.17 per cent. Acetylene, C_2H_2 , is also formed at temperatures exceeding 1800° . Acetylene is also formed when ethylene is passed through a hot glass tube: $3C_2H_4 = 2C_2H_2 + 2CH_4$; and also $C_2H_4 \rightarrow H_2 + C_2H_2$. Acetylene can be made by reducing carbon monoxide and by burning in coal gas (Fig. 270). To show the formation of acetylene when the Bunsen's burner "strikes back" and burns in the metal tube, arrange the apparatus shown in Fig. 270. A glass funnel is bent twice at right angles and connected with a cylinder containing ammoniacal cuprous

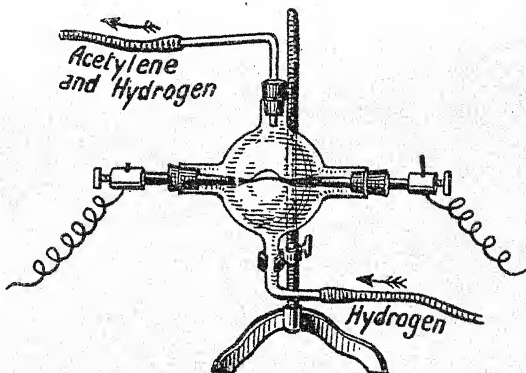


FIG. 269.—Berthelot's Synthesis of Acetylene.

chloride.¹ The Bunsen's burner is lighted at the pin-hole jet, and air is slowly aspirated through the apparatus. A red precipitate is formed in the cuprous chloride solution. This will be described later. The gas is most conveniently prepared by placing fragments of calcium carbide in a dry flask provided with a tap-funnel and delivery tube. The delivery tube *C* of Fig. 50

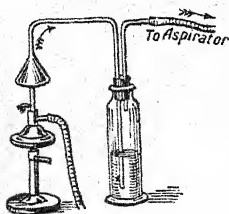


FIG. 270.—Acetylene from Coal Gas.

can be connected directly with the generating flask at *E*, Fig. 271. On gradually admitting water from the tap-funnel, a stream of acetylene is evolved. A modified generating flask is illustrated in Fig. 272. In this case the flask is filled with water, and the fragments of calcium carbide are added as required. The reaction between the water and the calcium carbide is represented by the equation: $\text{CaC}_2 + 2\text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{C}_2\text{H}_2$. The thermal value of the reaction is 29.1 Cals. The reaction is somewhat complex. The "balance-sheet" of the heat concerned in the reaction is usually given as:

<i>Heat liberated.</i>		<i>Heat absorbed.</i>	
	Cals.		Cals.
Formation Ca(OH)_2 . . .	160.1	Formation acetylene . . .	58.1
		Decomposition of water . . .	69.0
		Decomposition of carbide . . .	3.9
		Balance	29.1
Total	160.1	Total	160.1

The gas contains small traces of sulphur and phosphorus compounds, ammonia, etc., but it is usually pure enough for the experimental work described above. The gas can be purified from the most objectionable

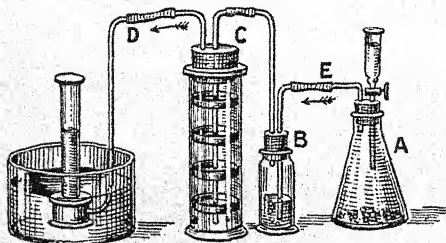


FIG. 271.—Preparation of Acetylene.



FIG. 272.—Rüchhoff's Flask.

impurities in the following manner. First pass the gas through a wash-bottle containing a solution of copper sulphate acidified with sulphuric

¹ Ammoniacal cuprous chloride is made by dissolving 10 grams of cupric oxide with 100 c.c. of concentrated hydrochloric acid, and the whole boiled for half an hour with an excess of metallic copper. Dilute the solution with an excess of water. Wash the precipitate twice by decantation with water. Dissolve the precipitate in a concentrated solution of ammonium chloride. If the solution is coloured brown, add a few drops of hydrochloric acid and some strips of metallic copper. For use, make the solution alkaline with a few drops of ammonia. If an acid solution of cuprous chloride is wanted, use hydrochloric acid instead of ammonia.

acid—*B*, Fig. 271. This removes ammonia, phosphorus, and sulphur compounds. The gas then passes through the tower *C* fitted with perforated shelves on which rest "chloride of lime," this removes the phosphorus compounds. This system of purification imitates some industrial plants for the preparation of large quantities of acetylene. In Siemens and Halske's process for the manufacture of hydrogen, superheated steam is allowed to act upon calcium carbide so that the formation of acetylene is avoided as much as possible. The main reaction is then: $\text{CaC}_2 + 5\text{H}_2\text{O} = \text{CaCO}_3 + \text{CO}_2 + 5\text{H}_2$.

Acetylene occurs in small quantities among the products of the distillation of coal gas.

Properties.—Acetylene is a colourless gas, which, when pure, has an ethereal odour which is not unpleasant. As usually prepared and purified the gas has traces of impurities which impart to the gas an offensive smell reminding one of garlic. Acetylene is rather lighter than air. 100 volumes of water at 0° dissolve 173 volumes of acetylene; and at 20° , 103 volumes of the gas. Alcohol dissolves about six times its own volume at ordinary temperatures; and the gas is absorbed by fuming sulphuric acid. Acetylene can be hydrogenized (reduced) to ethylene, C_2H_4 , and to ethane, C_2H_6 , if it be mixed with hydrogen, and passed over finely-divided nickel. Acetylene is poisonous and soon induces headache. The risk of poisoning with acetylene is much less than with carbon monoxide because acetylene is easily detected by its smell. The colour of the blood, in cases of acetylene poisoning, is said to be cherry red as with carbon monoxide poisoning, but the hæmoglobin is not affected in the same way. There is more hope of recovery with acetylene poisoning than with carbon monoxide poisoning.

Action of chlorine.—If a gas cylinder be partly filled with acetylene, and chlorine be allowed to pass into the cylinder bubble by bubble, the acetylene flashes as the chlorine enters, and deposits soot on the walls of the cylinder. Note that methane and ethylene when mixed with chlorine must be ignited before the soot is deposited. This experiment can be varied in an interesting manner by filling a cylinder about one-fifth full with a fresh solution of "chloride of lime," and then adding some hydrochloric acid. The cylinder will soon be filled with chlorine gas. Add a few pieces of calcium carbide the size of a pea. As soon as the acetylene comes in contact with the chlorine, it bursts into flame with the separation of large volumes of soot: $\text{C}_2\text{H}_2 + \text{Cl}_2 = 2\text{C} + 2\text{HCl}$.

Combustion.—Acetylene burns with a luminous smoky flame, but, like the other hydrocarbon gases, it extinguishes a lighted taper plunged into the gas. If acetylene be burned from a jet with a very fine aperture the flame is not smoky, but it is exceedingly luminous. In most acetylene burners the gas issues as two small jets so arranged that they strike against one another and produce a flat flame. Other holes are located so that air is drawn in and mixed with the gas as it rushes through the nozzle—*e.g.* the gas jet of an ordinary acetylene bicycle lamp. The great luminosity of the acetylene flame, coupled with the easy preparation of the gas from "carbide," has led to the extensive use of acetylene for bicycle lamps, houses, etc., where coal gas is not convenient; acetylene is also used to increase the luminosity of other inflammable gases. If the luminosity (candles per cubic foot) of methane be taken as one unit, the luminosity of ethylene is about 20; and acetylene, about 50. Acetylene is violently

explosive when mixed with $2\frac{1}{2}$ times its volume of oxygen. The gas cannot be safely stored under a greater pressure than 2 atmospheres—30 lbs. per square inch—because it is then liable to explode, violently by mere shock. One method of storing acetylene under pressure is to employ a solution of the gas in acetone, which, under a pressure of 12 atmospheres, dissolves 300 times its volume at ordinary temperatures. *Oxy-acetylene blowpipes* are used for welding pieces of iron and steel together under conditions where forge welding is impracticable. The flame is produced by burning a mixture of the two gases delivered into special blowpipes under pressure (the acetylene from an acetone solution of acetylene under pressure). The flame at the apex of the small central white cone has a temperature of about 3000° . At that point, the flame is almost entirely carbon monoxide surrounded by a jacket of hydrogen. The temperature at the apex of the flame is too high to allow the hydrogen to combine with the oxygen. The flame is therefore hot enough to melt iron and steel, and yet sufficiently reducing to protect the fused metal from oxidation while the welding is in progress. The oxy-acetylene blowpipe flame (2400°) is said to be hotter than the flame furnished by any other blowpipe—the oxy-hydrogen flame is about 2000° .

Action of heat.—Although stable at comparatively high temperatures—witness its formation in the electric arc (Fig. 269)—it is easily decomposed, with the separation of carbon, at lower temperatures. For instance, when heated in a glass tube to a temperature between 780° and 800° by means of an ordinary gas burner, Fig. 270, p. 854. The carbon which separates glows brightly owing to the heat developed during the decomposition of the acetylene: $C_2H_2 \rightarrow 2C + H_2 + 50 \text{ Cals.}$ These phenomena, (1) ready separation of carbon; and (2) the liberation of thermal energy which raises the temperature of the products of decomposition, appear to be related with the high luminosity of the acetylene flame. A commercial process for the manufacture of hydrogen is based upon this reaction. The separated carbon is employed as a high-grade lamp black. If acetylene be slowly passed through a glass tube at a dull red-heat, 500° to 600° , and then through a condenser, a few cubic centimetres of an oil will be obtained which, when distilled, furnishes a colourless volatile liquid which boils between 70° and 90° . It contains benzene as well as smaller quantities of anthracene, naphthalene, etc. Ethylene and methane appear to be formed at the same time. The benzene appears to be produced by the polymerization of the acetylene: $3C_2H_2 = C_6H_6$. Acetylene is also decomposed at still lower temperatures, 130° to 250° , in presence of finely divided metals—copper, iron, etc.

Acetylides.—The amount of acetylene in a mixture of different gases is determined by leading the mixture through an ammoniacal solution of cuprous chloride, when a reddish-brown precipitate of copper acetylide, usually represented by Blochmann's formula, $C_2H_2Cu_2O$, is formed. This, when dried, forms copper acetylide, C_2Cu_2 , which is explosive when heated between 50° and 90° , or subjected to percussion. The amount of acetylene is determined by filtering and washing the precipitate with ammoniacal water until the washing water is colourless. The precipitate is dissolved in hydrochloric acid, and the copper determined in the usual manner. Every gram of copper so obtained represents 0.17 gram of acetylene. The presence of acetylene in coal gas can be established

by sending a known volume of coal gas through the ammoniacal solution of cuprous chloride. An appreciable precipitate is formed in five or ten minutes. In the series of hydrogen compounds of nitrogen ranging from ammonia to azoimide or hydrazoic acid, a reduction in the proportion of hydrogen was attended by increased acidity; so of the carbon compounds, C_2H_6 , C_3H_4 , C_2H_2 , acetylene alone behaves like an acid because its hydrogen atoms can be replaced by a metal. The low solubility of acetylene in alkaline solutions shows that its acidic character is probably very feeble.

Composition.—The composition of acetylene has been established by methods similar to those employed for ethylene and methane. The results correspond with the molecule C_2H_2 , and the graphic formula $H-C\equiv C-H$.

§ 4. Hydrocarbons—Homology.

Between 250 and 300 compounds of carbon and hydrogen are known, and they can be arranged in a few series the members of which have many properties in common. The first member of each of the following three series is usually treated in inorganic chemistry, and all the series are discussed in organic chemistry.

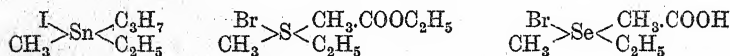
PARAFFIN SERIES.		OLEFINE SERIES.		ACETYLENE SERIES.	
	Boiling point.		Boiling point.		Boiling point.
Methane .	CH_4 -164°	—	—	—	—
Ethane .	C_2H_6 -93°	Ethylene .	C_2H_4 -105°	Acetylene .	C_2H_2 -83.6°
Propane .	C_3H_8 -45°	Propylene .	C_3H_6 -40°	Allylene .	C_3H_4 -23.5°
Butane .	C_4H_{10} 1°	Butylene .	C_4H_8 1°	Crotonylene .	C_4H_6 -27°
Pentane .	C_5H_{12} 35°	Amylene .	C_5H_{10} 39°	Valerylene .	C_5H_8 48°
Hexane .	C_6H_{14} 70°	Hexylene .	C_6H_{12} 69°	Hexoylene .	C_6H_{10} 80°
General formula C_nH_{2n+2}		General formula C_nH_{2n}		General formula C_nH_{2n-2}	

Any member of a series is represented by the general formula of its series. As recommended by C. Gerhardt (1843), each series is called an homologous series because there is a constant difference— CH_2 —between any one compound and the next higher or lower member so that all the compounds of the series appear to be proportional. In the paraffin series if n be less than five the hydrocarbon is gaseous at ordinary temperatures: from $n = 5$ to $n = 15$, liquid; and from $n = 16$ upwards, solid. The boiling or melting point usually rises with increasing values of n for each homologous series. The hydrocarbons of the first series are rather inert chemically, insoluble in water, and dissolve in one another in all proportions; they are excellent solvents for fats and similar substances; they burn readily, forming carbon dioxide and water; the gases explode when mixed with oxygen and ignited; the greater the carbon content the more luminous the flame. The terpene series of hydrocarbons— C_nH_{2n-4} —starts with valylene, C_5H_8 , and turpentine is the $C_{10}H_{16}$ member. The benzene or aromatic series— C_nH_{2n-6} —starts with benzene which is the C_6H_6 member; toluene is C_7H_8 ; and xylene, C_8H_{10} .

Organic chemists have developed a most extensive chemistry of the compounds of carbon which is based on the following hypotheses: (1) the valency of the carbon atom is always four; (2) the four affinity units are equivalent; (3) they are distributed in space in three dimensions, and

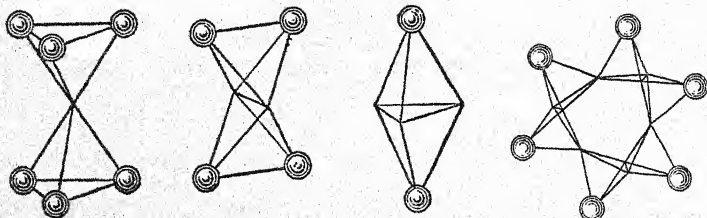
act in the direction of the axes of a tetrahedron; (4) the carbon atoms can unite with one another by means of one, two, or three affinity units and form chains of compounds, these chains may be open or closed. The number of carbon atoms so united may be very large. The closed chains usually contain three, four, five, six or seven carbon atoms in the ring, and in these chains (open or closed) some of the carbon atoms may be replaced by atoms of oxygen, nitrogen, sulphur, etc.

Constitution.—It is interesting to apply the hypothesis that the carbon atom behaves as if it were a regular tetrahedron to the 2-carbon member of each of the three series indicated above, viz. ethane, ethene or ethylene, ethine or acetylene, and benzene. The subjoined diagrams illustrate how the tetrahedra may be united in each case. Benzene is an illustration of a closed chain or six-carbon ring. It has been supposed that the decreasing stability of these compounds with an increase in the number of linking bonds is due to the *bending* or *straining* or *displacement* of the linking bonds from their most stable position—namely, that in which the linking bonds are directed from the centre towards the apices of a tetrahedron. When the four valencies of a carbon are satisfied by four different univalent groups, the resulting compound is usually optically active. This also applies to some other elements—tin, sulphur, and selenium. Thus:



represent optically active compounds. In organic chemistry there is some evidence for the belief that a series of compounds exist in which carbon is trivalent—e.g. triphenylmethane, $\text{C}(\text{C}_6\text{H}_5)_3$.

Polymerization.—It will be observed that each member of the olefine series might be regarded as an allotropic modification of the other.

Ethane— C_2H_6 .Ethylene— C_2H_4 .Acetylene— C_2H_2 .Benzene— C_6H_6 .

Chemical analyses show that they all have the same percentage composition, and all can be represented by the general formula $(\text{CH}_2)_n$. In the case of ethylene, $n = 2$; butylene, $n = 4$, etc. Similarly acetylene— C_2H_2 —has the same percentage composition as benzene— C_6H_6 . They differ from one another in their molecular weight. Polymerism is a word used to express the fact that two or more different compounds may have the same percentage composition but different molecular weights. Water is probably another example. We are almost certain that in steam most of the molecules are H_2O ; in liquid water probably H_3O_4 : and in ice the molecule is probably still more complex. The different polymeric modifications of

a compound may contain the same elements, but they appear to be associated with different proportions of available energy, *e.g.* $2C + 2H = C_2H_2 - 53.1 \text{ Cals}$; $6C + 6H = C_6H_6 - 82.8 \text{ Cals}$.

The student will have noticed that in naming compounds we usually place the more electropositive element first. Thus, from Table XXIV., hydrogen is more electropositive than oxygen, and accordingly water is said to be a hydrogen oxide, not an oxygen hydride.

§ 5. The Different Kinds of Chemical Action.

The immediate effect of affinity which a body exerts is always a combination; hence all the effects produced by chemical action are the consequences of the formation of some combination.—C. L. BERTHOLET (1803).

All chemical changes are reducible to union (identification) and division (differentiation).—T. S. HUNT (1861).

This is a convenient place to review the different kinds of chemical action so far considered:

I.—REACTIONS AMONG MOLECULES OF THE SAME SUBSTANCE.

1. **Isomeric changes.**—The atoms of the molecule undergo a rearrangement to form a new substance of the same composition as the old, but with different properties, as when ammonium cyanate, NH_4CNO , forms urea, $CO(NH_2)_2$.

2. **Polymerization.**—Two or more similar molecules may unite together to form a more complex molecule. *E.g.* three molecules of acetylene, C_2H_2 , may unite and form one molecule of benzene, C_6H_6 .

3. **Depolymerization.**—A complex molecule decomposes, producing two or more molecules of the same kind, *e.g.* the dissociation of nitrogen tetroxide, N_2O_4 , into two molecules of NO_2 .

4. **Condensation.**—When two (or more) molecules of a compound unite with the elimination of two (or more) atoms or radicals. *E.g.* when two molecules of sodium sulphite, in the presence of iodine, may each eliminate an atom of sodium, and then unite to form sodium dithionate, and the sodium atoms unite with the iodine to form sodium iodide.

5. **Decomposition or analytical reactions.**—A substance forms two or more different substances. *E.g.* mercuric oxide furnishes mercury and oxygen.

II.—REACTIONS BETWEEN MOLECULES OF DIFFERENT COMPOUNDS.

6. **Combination or synthetical reactions.**—Two or more different substances unite to produce another substance. *E.g.* zinc and oxygen unite to form zinc oxide; calcium oxide and carbon dioxide form calcium carbonate; ethylene and chlorine form ethylene dichloride (addition product).

7. **Metathesis or exchange.**—Two or more substances interact to form two or more new substances. Metathesis, from the Greek *μετά* (meta), beyond; *τιθέω* (titheo), I place. The transposition may involve:

(a) Simple displacement, replacement, or substitution of one radicle or element for another. *E.g.* metallic iron with copper sulphate

forms metallic copper and ferrous sulphate; zinc and hydrochloric acid give zinc chloride and hydrogen.

- (b) Double decomposition or mutual exchange such as occurs during hydrolysis, neutralization, etc. *E.g.* sodium hydroxide and hydrochloric acid give water and sodium chloride; silver nitrate and sodium chloride give silver chloride and sodium nitrate.

§ 6. Natural Gas, Petroleum and Related Products.

In the inflammable gas issuing from the earth in the vicinity of petroliferous formations, Nature, with profuse liberality, has provided a clean, serviceable fuel for the use of the engineer.—ANON.

The firedamp of the miner is a natural gas which has been known from a very remote period. The jets of gas which issue from fissures in the earth in many parts of the world are constant or intermittent and the gas is of a similar nature. The perpetual fires in the district about Baku on the Caspian Sea were objects of awe and wonder. The flames were maintained by an invisible mysterious agent. The unknown was taken to be a supernatural agent, and the blazing jets were considered by the untutored, credulous, and superstitious ancient dwellers in those regions to be a convincing and direct proof of the presence of a deity. Altars and temples were built on the hills for the worship of the unknown god who was called "the soul-of-the-world," or "the spirit of fire." To-day we know that the perpetual fires are merely jets of hydrocarbon issuing from fissures communicating with a subterranean bed of coal, shale, or oil reservoir. The hydrocarbon vapours being inflammable, burn in contact with the air. The emission of natural gas from the earth at great pressures has been utilized by chemical engineers; and in the United States, where immense supplies of natural gas are available, the gas is distributed to consumers in towns, and numerous industrial processes are carried on with its aid as a fuel. The gas is usually ready for immediate consumption, and is somewhat similar to and sometimes superior to the gas prepared in producers, gasworks, etc.

Petroleum.—Crude petroleum—also called *rock oil*—is a thick viscid liquid varying in colour from straw-yellow to greenish-black, and most varieties show a greenish fluorescence by reflected light. Petroleum is a complex mixture of many hydrocarbons belonging principally to the paraffin series along with small quantities of nitrogen and sulphur compounds. Petroleum occurs in the Baku district in Russia; in California, Colorado, Indiana, Kansas, Kentucky, New York, Ohio, and Texas in the United States; in Mexico, Canada, India, Egypt, South and West Africa, Peru, Trinidad, Barbadoes, Borneo, Burmah, Australia, New Zealand, etc.

The oil apparently occurs underground, and in some places it issues from the earth without man's assistance. It is usually necessary to "bore" through the overlying strata and insert a pipe in the "oil basin." When first "tapped," the oil often "shoots" out of the "well," but the velocity of the stream gradually subsides,¹ and, after a time, the oil is "pumped" to the surface; conveyed by pipes to a central reservoir for

¹ When the Lucas oil-well (Beaumont, Texas) was first "tapped" in 1901, a six-inch stream of oil is said to have spouted 160 feet high for 9 days at the rate of 75,000 barrels per day.

storage; and afterwards distributed. Fig. 273 is a diagrammatic sketch through the strata of an oilfield, and it is intended to give a rough idea of the way the oil and gas are associated in some oilfields. The gas and oil here collect near the summit of an anticlinal (concave downwards) fold capped by an impervious rock—called “cap-rock.” If the gas is under pressure, it is easy to understand, from Fig. 273, how the oil from the “bore” shown in the diagram would be expelled with some violence until the pressure is relieved. Saline water is usually associated with oil and gas. In some cases the oil is obtained from a horizontal stratum of porous sandstone or limestone saturated with oil.

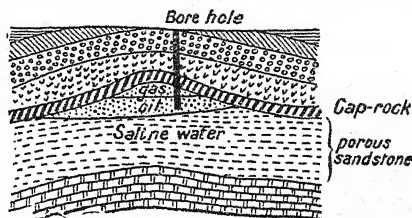


Fig. 273.—Geological Section of Oil-well (Diagrammatic).

Origin of petroleum.—The origin of petroleum is unknown.

Some argue: (1) that petroleum is a product of the slow distillation of animal or vegetable products—at high or at low temperatures. It is assumed in some cases that processes analogous with the manufacture of coal gas are being performed on a colossal scale in the bowels of the earth. While the commercial process for making coal-gas gives a preponderance of hydrocarbons of the benzene series, natural petroleum usually have the paraffin series predominant. If, however, coals be distilled at a low temperature under reduced pressure, the products approach very nearly to those obtained with natural petroleum. Thus, by distilling coal in vacuo at 450°, Pictet and Bouvier (1913) obtained a tar which yielded products similar to those obtained with Canadian petroleum. Hence, it is inferred that petroleum is one of Nature's by-products from the low temperature distillation of coal. The gas obtained by such a process contains more carbon dioxide and monoxide than is found in natural gas, but this would be absorbed from the gas by alkaline subterranean waters produced by the weathering of rocks. Carbon monoxide in contact with hydrogen is oxidized to the dioxide at temperatures as low as 250° in the presence of ferric or cupric oxide. Hence both carbon monoxide and dioxide would disappear. Others—Mendeléeff, Berthelot—argue (2) that the oils and gas are produced by the interaction of water with metallic carbides—say, iron carbide—at great depths. (3) P. Sabatier has shown that acetylene in contact with nickel at different temperatures and with varying proportions of hydrogen, furnish complex hydrocarbons which in every respect resemble natural petroleum. The claims of these and other rival theories cannot be established in ignorance of the conditions which prevail in the locality where the petroleum were formed. Possibly each theory may be valid in some special district.

Refining.—Crude petroleum is used as a fuel in many industries—metallurgy, locomotives, fire-engines, steamships, etc. A great deal of petroleum is purified or refined. The treatment of petroleum oil and its products is a vast industry. Over 200 different commercial products are derived from the purification and refining of petroleum. The crude petroleum is placed in a retort—“still”—connected with condensing

tubes and receiving tanks. The temperature is gradually raised. At first, the lighter substances are volatilized and condensed in suitable receivers. The receivers are changed when the specific gravity of the distillate has attained a certain value, or when the temperature of the retort has been raised sufficiently high. The chief fractions are :

TABLE LIII.—PRODUCTS OF THE DISTILLATION OF PETROLEUM.

Fraction.	Chief contents.	Approximate boiling point.	Uses.
Cymogene	C_4H_{10}	About 0°	Artificial cold.
Rhigolene	C_4H_{10} to C_5H_{12}	16°	Local anæsthetic by freezing.
Petroleum ether	C_5H_{12} to C_6H_{14}	$50-60^\circ$	Solvent; fuel.
Gasoline; petrol	C_6H_{14} to C_7H_{16}	$70-90^\circ$	Solvent; fuel.
Ligroin; naphtha	C_7H_{16} to C_8H_{18}	$90-120^\circ$	Solvent; fuel.
Benzine (not benzene); benzoline	C_8H_{18} to C_9H_{20}	$110-140^\circ$	Solvent; substitute paint oil.
Kerosene; photogene; paraffin oil	$C_{10}H_{22}$ to $C_{17}H_{36}$	$150-300^\circ$	Fuel; illuminant.

The residue in the retort is transferred to another still, and further heated to a high temperature. It furnishes lubricating oils; vaseline ($C_{19}H_{40}$ to $C_{21}H_{44}$) used for ointments, etc.; paraffin ($C_{21}H_{44}$ to $C_{32}H_{66}$, melting between 45° and 76°) used for candles, insulating, etc. The residue in the retort is mainly coke. The products may be still further purified. For instance, kerosene is washed with sulphuric acid, and then with caustic soda or sodium carbonate and water to get rid of tarry matters and "semi-solid" hydrocarbons which might choke the wicks of lamps; and afterwards redistilled to remove oils of low flash-point (p. 909) which might cause an explosion when the oil is used as an illuminant. Sulphur compounds can be removed, to a certain extent, by treatment with cupric oxide. The methods for removing sulphur from these oils are mainly trade secrets.

There is no record where bitumen was first discovered. It is mentioned many times in the Bible—e.g. Genesis ix. 3—and Strabo said that it was used by the Egyptians in emblaming the dead. Centuries before the Christian era, Baku was a resort for the fire-worshippers of Persia and India, and not until 1880 were pilgrimages stopped by the Russian Government.

Ozokerite, or ozocerite, is a wax-like, dark yellow or brown, native paraffin with a greenish opalescence. It is found associated with petroleum in the sandstone of Galicia, where it is extensively mined. Small quantities are found in other places. Ozokerite when bleached and purified furnishes ceresine—used as a substitute for beeswax, for making ointments, candles, and bottles for storing hydrofluoric acid. Ozokerite and asphaltum appear to be residues left after the natural distillation of petroleum. Asphaltum or mineral pitch occurs in quantity in the "pitch lake" of Trinidad. It may occur as a soft brownish-black substance, or as a black solid. It is a mixture of different hydrocarbons and resembles artificial

asphaltum or pitch obtained by the distillation of coal tar. It is used for making pavements, waterproofing materials, etc.

Oil shales.—These shales are associated with the sandstones, clays, and limestones of the "calciferous sandstone series" in Mid and West Lothian and other parts of Scotland. The oil shales of New South Wales are also called "kerosene shales." Good oil shales can often be lighted with a match, when they burn with a steady flame resembling a candle. When heated to dull redness in vertical retorts, they furnish gas and a liquid distillate which separates into two layers—the lower aqueous layer contains ammonium compounds; the upper layer has a greenish-brown colour and it contains oil and tar. This layer closely resembles petroleum, and gives similar products on fractional distillation. The distillation of oil shales is facilitated by blowing low-pressure steam into the retorts. Scotch shales furnish from 18 to 50 gallons of crude oil per ton; New South Wales shales are said to yield up to 100 or 150 gallons of crude oil per ton. The shale-oil of Pumpherston (Scotland) was distilled five times and scoured four times with sodium hydroxide, and with sulphuric acid before it furnished the "clear water-white" oil. *Dry distillation* or destructive distillation are terms applied to the decomposition of a substance in a closed vessel so as to obtain the volatile products.

§ 7. The Calorific Power of Fuels.

Thermal energy, heat, is largely employed for domestic and industrial purposes, and a very large proportion of the mechanical and electrical energy employed in the industrial world is really derived from the combustion of carbon in the form of coal. In other words, during combustion, the chemical energy "stored" up in the fuel is degraded in the form of heat energy, which in turn is transformed into mechanical and also into electrical energy. There is unfortunately a tremendous percentage loss in the transformation, and one of the most important problems confronting the chemical engineer is to reduce this loss to a minimum. The amount of fuel necessary for inaugurating chemical reactions in metallurgy is one of the most important controlling factors in the commercial success of many processes. There is also quite a host of commercial problems which depend on the utilization of chemical energy, so that the energetics of chemical processes should be studied with as much diligence as the weight and volume relations of the substances concerned in chemical reactions.

The commercial value of coal, to a large extent, is determined by its heat of combustion, and, consequently, many prefer to purchase coal by a scale based on its heating power, not merely on its price per ton. Other things being equal, a coal bought at 10s. per ton might prove much cheaper than coal at, say, 9s. 10d. per ton, because the heating power, i.e. the available chemical energy, of the former might be greater. In commercial work, the amount of heat furnished by the combustion of unit weight (pound, gram, or kilogram) of the fuel is called the **calorific power** of the fuel. The unit of heat may be the amount of heat required to raise the temperature of one pound of water 1° C.—this is called the **pound calorie**; if 1° F. is used, the so-called **British thermal unit**—B.T.U.—is obtained. Kilogram and gram-calories are also used.

The heat of combustion of carbon (charcoal) is 96,980 cal. (p. 239).

This means that 12 grams of carbon will furnish, on combustion to carbon dioxide, 96,980 calories. Hence one gram of carbon will furnish $96,980 \div 12 = 8080$ cals. This number, 8080 cals., is taken to represent the calorific power of the carbon. The calorific power of a few important constituents of fuel are :

	calories.
Carbon to CO_2	8,080
Carbon to CO	2,400
Carbon monoxide	5,680
Hydrogen (to liquid water)	29,300
Methane, CH_4 (to liquid water)	11,850
Ethylene, C_2H_4 (to liquid water)	10,460
Acetylene, C_2H_2 (to liquid water)	11,500

The calorific power of coal determined in a bomb calorimeter is not very far removed from that calculated from the ultimate composition of the coal on the assumption that the oxygen in the coal will render one-eighth of its own weight of hydrogen useless, so far as the development of heat is concerned.

EXAMPLE.—A sample of coal furnished, on analysis, 73 per cent. of carbon ; 6.0 per cent. of hydrogen ; and 16 per cent. of oxygen. The other constituents were non-combustibles. What is the calculated calorific power of the coal ? The analysis means that 1 lb. of the coal contains 0.73 lb. of carbon ; 0.06 lb. of hydrogen ; and 0.16 lb. of oxygen : $\frac{1}{8}$ of 0.16 is 0.02, hence, 0.04 lb. of hydrogen is available for heating purposes. The carbon furnishes $8080 \times 0.73 = 5898.4$ cals. and the hydrogen $0.04 \times 29,300 = 1172$ cals. Adding these two results, the calculated calorific power of the coal is 8070.4 cals.

The calorific power of gaseous fuels—producer gas, water gas, etc.—can be determined in a similar manner from the table which precedes. There is a small complication in that the analysis of the gases is usually represented by volume. The method indicated on p. 85 is employed to convert the volumes into weights. The calorific power refers to the thermal value of unit weight (pound, gram, or kilogram). With gases it is more convenient to express the result as the thermal value of 1000 cubic feet of gas.

EXAMPLE.—It is required to find the heat of combustion of 1000 c. ft. of a sample of coal gas which furnished, on analysis : hydrogen, 48 per cent. ; carbon monoxide, 8 ; methane, 36 ; ethylene, 3.8 ; nitrogen, 4.2 per cent.

	Percentage composition of gas.	Weight of 1 c. ft. in lbs.	Total weight per lb.	Percentage weight.	Weight per lb. of gas.
Hydrogen .	48.0	0.0056	0.269	8.6	0.086
Carbon monoxide . .	8.0	0.0871	0.625	20.0	0.200
Methane . .	36.0	0.6447	1.609	51.4	0.514
Ethylene . .	3.8	0.0784	0.298	9.5	0.095
Nitrogen . .	4.2	0.0784	0.329	10.5	0.105
	100.0	—	3.130	100.0	1.000

Thus, 100 c. ft. of the gas weighs 3.13 lbs.

Again,

Combustible constituents.	Weight per lb. of coal gas.	Calorific power per lb.	Calorific power, calories.
Hydrogen	0.086	29,300	2,517
Carbon monoxide	0.200	2,400	479
Methane	0.514	12,000	6,169
Ethylene	0.095	10,400	990
Total			10,155

100 c. ft. of the gas weighs 31.3 lbs., and 1 lb. of the gas furnishes 10,155 cals., consequently, 1000 c. ft. will furnish $10,155 \times 31.3 = 317,900$ cals.

The student, after solving the problems indicated in what precedes, should have no difficulty in calculating the amount of air required for the combustion; in calculating the composition of the products of combustion; and conversely, as is sometimes needed in industrial work, in calculating the amount of air in excess of that required for complete combustion given the amount of carbon dioxide and oxygen in the flue gases (products of combustion).

A popular way of showing the great amount of energy in coal is to calculate that the heat evolved during the complete combustion of a pound of coal is sufficient to raise 10,000,000 pounds a foot high. In attempting to convert the chemical into mechanical energy, about 90 per cent. is lost. The best of steam engines can utilize but one-eighth to one-tenth of the available energy.

§ 3. The Temperature of Combustion.

Although the heat of combustion is constant for a definite substance, the actual temperature attained by the combustion is dependent upon a number of factors. One pound of carbon furnishes 8080 (pound) cals. The combustion of one pound of carbon gives $3\frac{2}{3}$ lbs. of carbon dioxide; the specific heat of carbon dioxide is 0.216. From the well-known formula: Quantity of heat is equivalent to the mass of substance heated multiplied by the specific heat of the substances heated multiplied by the rise of temperature, we get:

$$8080 = 3\frac{2}{3} \times 0.216 \times \text{Rise of temperature.}$$

Hence the rise of temperature is $10,180^\circ$. We have assumed that the carbon was heated in oxygen, and the heat of combustion is spent in raising the temperature of the products of combustion. If the carbon were burnt in air, the $2\frac{2}{3}$ lbs. of oxygen required for the complete combustion of carbon would be accompanied by 8.9 lbs. of nitrogen. If nitrogen be present as well as oxygen, part of the heat will be spent in raising the temperature of the nitrogen. Nitrogen has a specific heat 0.244. Hence:

$$8080 = (3\frac{2}{3} \times 0.216 + 8.9 \times 0.244) \times \text{Rise of temperature}$$

or the rise of temperature will be 2733° . If an excess of air be present, the temperature will be still further reduced. Hence the calculation of

the heat of combustion of a substance requires a knowledge of the composition of the mixture heated, the specific heat of the products of combustion, etc. If the combustion be slow, some of the heat may be lost by conduction, radiation, etc. Then again, the specific heat of gases increases with rising temperatures so that the specific heat of a gas determined at low temperatures, say 100° , is not the same as the specific heat of the gas at, say, 1000° . Hence, calculations of the temperature of combustion, made in ignorance of these factors, are not of much practical value, although they are sometimes useful for purposes of comparison.

EXAMPLE.—What is the heat of combustion of methane in oxygen and in air when the calorific power is 12,000? Given the specific heat of steam, 0.480 ; nitrogen, 0.244 ; and carbon dioxide, 0.216 . Ansr. Nearly 7160° in oxygen, and 2430° in air. Hint. 2.75 lbs. of carbon dioxide, 2.25 lbs. of steam, and 13.4 lbs. of nitrogen are concerned in the combustion of 1 lb. of methane.

It is also necessary to remember that when the intensity factor of thermal energy liberated during combustion gets very high, the products of combustion may be "unburnt" by dissociation, for the consumption of energy by the dissociating gases cools the system. This dissociation reverses the process of combustion of fuel with oxygen. In this sense it may be said that the rise of temperature during combustion counteracts the affinity of the fuel for oxygen, and tends to reduce the temperature of combustion below that computed on the assumption that combustion is complete. Given the amount of dissociation it is easy to make the proper allowance. It also follows that there is a limit to the temperature attainable by the combustion of a given fuel.

§ 9. Gunpowder.

Of the mischievous composition and diabolical abuse of gunpowder, much might be written, but because the present world taketh only delight in shedding innocent blood, and cannot endure that unrighteous things should be reprov'd, and good things praised, therefore it is best to be silent.—J. R. GLAUBER (1658).

If potassium nitrate be mixed with powdered charcoal, and heated, the two materials react with explosive violence, forming potassium carbonate, nitrogen, and carbon dioxide: $4\text{KNO}_3 + 5\text{C} = 2\text{K}_2\text{CO}_3 + 2\text{N}_2 + 3\text{CO}_2$. The volume of the gases produced is so much greater than that of the original volume of the mixed solids that if the powder be ignited in a closed space, the expanding gases give the mixture the propelling, tearing, and splitting powers characteristic of explosives. It was soon found that the explosive effect is greater if the nitre and charcoal be mixed with sulphur, so that instead of solid potassium carbonate a residue of solid potassium sulphide is obtained; though side reactions lead to the formation of other products. The mixture is called gunpowder. Theoretically the main reaction is represented: $4\text{KNO}_3 + \text{S}_2 + 6\text{C} = 2\text{K}_2\text{S} + 2\text{N}_2 + 6\text{CO}_2$; though side reactions lead to the formation of other products. As an exercise on the methods of calculation indicated on the preceding pages, we can compute the approximate pressure developed during the explosion of gunpowder.

Problem.—To calculate the approximate pressure developed during the explosion of gunpowder in a closed vessel. For ease in calculation, take the atomic weights

C, 12; S, 32; O, 16; N, 14; K, 39. It follows that the theoretical mixture will contain 404 grams of potassium nitrate; 64 grams of sulphur; and 72 grams of carbon. Otherwise expressed, 75 per cent. of nitre; 12 per cent. of sulphur; and 13 per cent. of charcoal. This very nearly represents the average composition of gunpowder which is usually stated to be: nitre, 75; charcoal, 14; sulphur, 10; water, 1. The theoretical equation also shows that 220 grams of potassium sulphide; 56, nitrogen; and 264 of carbon dioxide are formed. Otherwise expressed, gunpowder on explosion furnishes 59 per cent. of gas; or one gram of gunpowder, at 0°, and 760 mm. pressure, furnishes 247.3 c.c. of carbon dioxide and 79 c.c. of nitrogen; in all, 327 c.c. of gas consisting of 0.49 gram of carbon dioxide; 0.10 gram of nitrogen; and 0.41 gram of potassium sulphide. Again, one gram of an average gunpowder occupies 0.9 c.c. The surface exposed by one c.c. is 6 square cm., hence 0.9 c.c. will expose 5.4 square cm. But if 0.9 c.c. of gunpowder be confined at atmospheric pressure it follows that 327 c.c. will be confined under $327 \div 0.9 = 363.2$ atmospheres pressure; or, if one gram of gunpowder at 0° be confined in a closed space and exploded, it furnishes sufficient gas to give $363.2 \div 5.4 = 67.3$ atmospheres pressure per square cm.

The reaction indicated above is exothermal, and much heat is developed. The rise of temperature will cause the gas to expand with an ever-increasing pressure. One gram of carbon in burning to carbon dioxide develops 8080 cal. Hence 0.13 gram of carbon will furnish 1050 cal. Assuming that the specific heat, that is the amount of heat required to raise the temperature of one gram of the substance 1°, is constant; and that the specific heat of carbon dioxide is 0.22; of potassium sulphide, 0.4; and of nitrogen, 0.24; remembering also that the quantity of heat Q is equal to the product of the weight of the substance heated, w , the rise of temperature x , and the specific heat s , we have $Q = wsx$; or,

$$1050 = \{(0.49 \times 0.22) + (0.1 \times 0.24) + (0.41 \times 0.4)\}x; \text{ or, } x = 3540^\circ$$

This means that the combustion of one gram of gunpowder will give sufficient heat to raise the temperature of the products of combustion 3540°. If 326 c.c. of gas be heated 3540°, the pressure corresponds with 880 atmospheres per square cm. Experiment shows that the observed pressure is but half that indicated by this theoretical discussion. The difference is due to several disturbing effects. (1) The analysis of the gaseous products of combustion shows that side reactions must also be in progress, for part of the oxygen forms K_2SO_4 , some of the carbon burns to carbon monoxide; some of the nitrogen to nitric oxide; some hydrogen and hydrogen sulphide are produced by the decomposition of the water present in gunpowder; and some of the gunpowder remains unburnt; (2) the apparatus in which the test is made is slightly elastic, and this interferes with the accurate measurement of the pressure; and (3) the specific heat of the gas increases appreciably with rise of temperature.

Gunpowder appears to have been introduced into Europe from the Saracens, and it is said to have been first used by the English at the battle of Werewater in 1327, and with memorable effect by Edward III. at the battle of Crecy (1346). Roger Bacon is sometimes credited with the discovery, but there is every reason to suppose that he obtained his knowledge from an Arabian source.

Modern chemistry has furnished man with explosive compounds—nitroglycerol and gun-cotton—which, under right conditions, do their work by decomposing with the evolution of a large volume of gas so rapidly that atmospheric air acts almost as if it were a resistant solid. *Nitroglycerol* is made by treating glycerol with concentrated nitric acid, and *gun-cotton* by treating cotton with the same acid. The violence of these compounds is tempered by mixing them with inert or less active substances. Thus, nitroglycerol absorbed by kieselguhr is *dynamite*: and *smokeless* powders are mixtures of gun-cotton and various "secret" substances. These explosives—excluding dynamite—give no solid product, and no "smoke" although a cloud of condensed water-vapour may be produced.

§ 10. Coal Gas.

The manufacture of coal gas for purposes of illumination is one of the most striking instances of science enlisted in the divine cause of civilization. Once a luxury or convenience, it is now almost a necessity.—ANON.

Stephen Hales, in 1726, subjected coal to distillation, and succeeded in preparing an inflammable gas resembling natural gas; but the possibility of manufacturing an artificial source of illumination from coal gas was not seriously considered until near the end of the 18th century. W. Murdock distilled coal in an iron retort, and lighted his house with the gas so manufactured. P. Leoss, in 1786, lighted his house with the gas obtained by the distillation of wood, and, in 1800, proposed to light the streets of Paris. The project fell through because of the feeble luminosity of wood gas. After W. Murdock's essay, several houses and factories were lighted with gas, but there were many complaints of the injurious effects of the products of combustion. No attempt had then been made to purify the gas. Coal gas had a great struggle for existence in its early days, but this only stimulated inventors to improved processes of manufacture and application. The early gas, it was claimed, produced "acids, stinks, and smokes," and there was a constant risk of "explosions, sore eyes, and blackened countenances." The streets of London were lighted with coal gas in 1812, and those of Paris in 1815.

When coal is heated in closed vessels to about 400°, it is carbonized, and a comparatively small quantity of gaseous, and a relatively large quantity of liquid, products are obtained. The hydrocarbon gases consist mainly of members of the methane and ethylene series. Benzene, acetylene, and hydrogen are generally present. If the temperature of distillation be raised, the quantity of liquid products decrease, and the quantity of gaseous products increase. In other words, more gas and less tar is obtained. The gas obtained by the high temperature distillation has less illuminating power. This is illustrated by the following table:

TABLE LIV.—EFFECT OF TEMPERATURE OF FORMATION ON THE ILLUMINATING POWER OF COAL GAS.

Approximate temperature.	Volume of gas—c. ft.	Illuminating power—candles.
420°	1,400	—
700°	8,250	20.5
900°	9,690	17.8
1000°	10,820	16.7
1200°	12,000	15.6

This decrease in the illuminating power is due to the decomposition of the hydrocarbons into simpler compounds—hydrogen and gas carbon. The illuminating power of the gas depends upon the proportion of "total hydrocarbons" present, and not on any single one. In modern practice there is a tendency to raise the temperature of distillation, thus sacrificing quality (*i.e.* illuminating power) for quantity (*i.e.* cubic feet per ton of coal). The gas obtained from coal is not uniform, but varies in com-

position with the nature of the coal, the temperature of decomposition, etc. The tar, carbon dioxide, sulphur and ammonia compounds, etc.—produced during the distillation of the coal—are removed from the gas, and finally, when the distillation is conducted in the neighbourhood of 1000°, the purified coal gas contains approximately :

Hydrogen.	Methane.	Ethylene.	Carbon monoxide.	Carbon dioxide.	Nitrogen.	Oxygen.
49	35	4	4	$\frac{1}{2}$	4	$\frac{1}{2}$ per cent.

The approximate proportions of the by-products produced at the same time, are, per ton of coal :

	lbs.	per cent.
Coal gas (10,000 c. ft.)	380	17.0
Tar	115	5.1
Gas liquor (without water from scrubbers)	177	7.9
Coke	1568	70.0

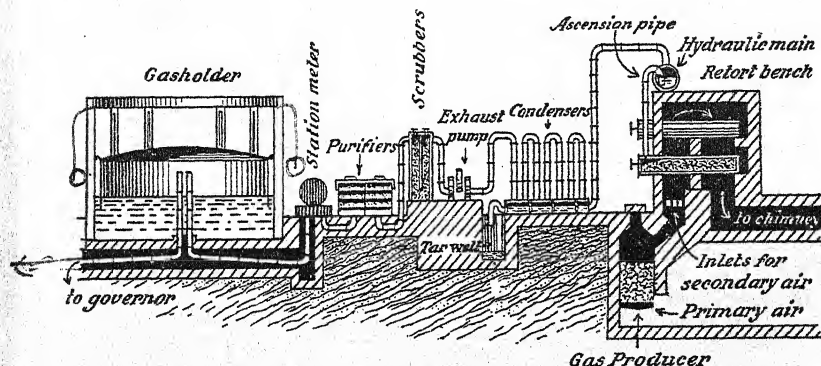


FIG. 274.—Coal Gas Works (Diagrammatic Section).

There are differences in detail in the manufacture at different gasworks. The following brief description assisted by the diagrammatic sketch, Fig. 274, will give a rough idea of the process of manufacture.

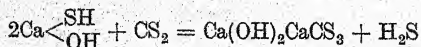
1. The retorts.—The coal is distilled in U-shaped fireclay retorts 6 to 8 feet long. The retort may be set horizontally, vertically, or slanting. On small works, the retorts are closed at one end; and in larger works, open at both ends. The retort is fixed to the iron furnace front which is provided with a door, and connected with a vertical exit pipe—the *ascension pipe*—for the escape of the volatile products. The retorts are generally arranged in tiers of 5, 7, or 9, so that they can be heated by a single furnace. The charge—about 200 lbs. of coal—is distributed evenly over the bottom of each retort and the mouth of the retort luted airtight. The retorts are heated by the furnace which will be discussed later. The volatile products pass from the retort into the *hydraulic main* which contains condensed coal tar and water. This main runs horizontally over the front of the bench of retorts, and all the retorts discharge into it. The gas is here partly cooled, and some tar and water

are condensed from the hot gas. The pipe leading the gas from the retort dips into the liquid in the hydraulic main, and so prevents the gas from passing back—"back lash"—when the retorts are recharged. The liquid in the hydraulic main is kept at a constant level by leading any excess into the *tar well*. In from four to six hours, the distillation will be complete. The coke is pushed or raked from the retort, and quenched with water to prevent further combustion. Another charge of coal is quickly introduced into the retort. The residual coke may be used for heating the retorts, and part is sold.

2. The condensers.—The hot impure gases pass from the hydraulic main into a series of iron pipes, several hundred feet long—the condensers—connected so that the gas must pass through the entire length of the pipe. The gas is here cooled still further, and more tar is condensed and run to the tar well. The condensed liquid in the tar well separates into two layers—the lower layer is *gas tar*; and the upper aqueous solution containing ammonia and ammonium salts is the *gas liquor*. The gas is drawn from the hydraulic main through the condensers by means of an *exhaust pump* which reduces the pressure in the retort, and also regulates the pressure of the gas sent along to be still further purified.

3. The scrubbers.—In modern works all the tar is removed from the gas in the condensing plant, but the gas still contains sulphur compounds, carbon dioxide, some ammonia, and possibly some tar. In one form of scrubber, a tower is filled with trays charged with coke or pebbles. The tower has a partition so that the gas flows down one side of the tower and up the other. A spray of water trickles down the coke. The gas in passing through the coke is broken up into small bubbles and washed free from ammonium compounds by the water. The water is drawn off intermittently at the base of the tower and mixed with the gas liquor from the tar well. The ammonia is recovered as a by-product.

4. The purifiers.—Some of the hydrogen sulphide and carbon dioxide in the gas combine with the ammonia and are removed in the scrubbers. The gas still contains sulphur compounds. If these were not removed, the burning gas would form sulphur dioxide¹ which is objectionable. The object of the purification is to remove the sulphur compounds and the carbon dioxide. The gas leaving the scrubbers is directed into a series of low rectangular iron tanks—the purifiers—fitted with horizontal shelves or grids. The shelves are loosely packed with a layer of slightly damped slaked lime—say, six inches deep. The lime removes hydrogen sulphide and carbon dioxide. A mixture of calcium sulphide— CaSH.OH , or Ca(SH)_2 —and calcium carbonate is formed. The calcium sulphide may absorb some carbon disulphide:



When the lime is spent or fouled it is called *gas lime*, or *spent lime*. To make sure that all the sulphur compounds are removed, the gas is generally passed through another purifier containing ferric hydroxide ("bog iron ore")— $\text{Fe}_2\text{O}_3.\text{H}_2\text{O}$. The ferric hydroxide forms ferric

¹ There is no experimental evidence to show that the sulphur forms sulphuric acid in a moist atmosphere unless the temperature be so low that the water is condensed to a liquid (cf. p. 502).

sulphide: $\text{Fe}_2\text{O}_3\text{H}_2\text{O} + 3\text{H}_2\text{S} = \text{Fe}_2\text{S}_3 + 4\text{H}_2\text{O}$; or ferrous sulphide and free sulphur: $\text{Fe}_2\text{O}_3\text{H}_2\text{O} + 3\text{H}_2\text{S} = 2\text{FeS} + \text{S} + 4\text{H}_2\text{O}$.

When the mixture is fouled it is placed in a heap for about twenty-four hours,¹ and then spread out in layers twelve inches deep. The layers are turned over repeatedly to expose fresh surfaces to the action of the air. The black iron sulphides are oxidized by exposure to the air, and free sulphur separate: $2\text{Fe}_2\text{S}_3 + 3\text{O}_2 = 2\text{Fe}_2\text{O}_3 + 6\text{S}$; $4\text{FeS} + 3\text{O}_2 = 2\text{Fe}_2\text{O}_3 + 4\text{S}$; the net result is that the hydrogen sulphide of the gas is converted into free sulphur, and the ferric oxide is revived ready to be used again. This alternate fouling and oxidizing of the "iron" is repeated about sixteen times when so much sulphur accumulates—55 per cent.—that it is no longer economical to use the oxide again. The *spent oxide* is sold to the manufacturer of sulphuric acid, and used as a source of sulphur.

5. The gas holder.—The purified gas next passes through a large meter—*station meter*—which records its volume. The gas holder is an enormous cylindrical iron tank which floats in a cistern of water, and rises or falls as gas enters or leaves. The cylinder is so weighted that the gas can be expelled from it at the necessary pressure. From the gas holder, the gas passes to the *governor*, where its pressure is reduced and regulated so as to give a supply of gas at the necessary pressure.

6. The by-products.—(1) *Coke* is a valuable fuel and finds a ready sale. (2) *Gas carbon* is a hard dense deposit of almost pure carbon which gradually collects on the inside of the retort. It is a good conductor of electricity, and is used for the manufacture of carbon rods for electric lighting, and of plates for galvanic batteries. (3) *Gas lime* is used for agricultural purposes. (4) *Tar*—gas tar, coal tar—is a black viscid foul-smelling liquid used as a protective paint for preserving timber; making tarred paper, waterproofing masonry, etc. Tar is a mixture of many "organic substances" which are separated by distillation at different temperatures. It furnishes carbolic and creosotic oils, benzene, naphthalene, anthracene, etc. Benzene is a most interesting compound discovered by M. Faraday; as A. W. Hofmann has said, its history and uses would fill a volume. It is capable of undergoing an endless number of Proteus-like transformations, and it has given rise to new and important branches of industry. It is used in the manufacture of aniline dyestuffs, flavouring essences, perfumes, oils, etc. The residue in the retort is "pitch." Asphalt is a solution of pitch in heavy tar oils, and is used in making hard pavements, varnish, etc. (5) *Ammonia*. The ammoniacal liquid is boiled with milk of lime and the expelled ammonia is mixed with sulphuric acid. The tarry matters are separated, and the solution of ammonium sulphate is evaporated and crystallized for the market.

§ 11. Producer Gas.

Reviewing the process just outlined one cannot but admire the results of the work of the chemist and gas-engineer. The marvellous ingenuity displayed in the construction of apparatus; the great skill exercised in

¹ When fouled for the first time the mixture may absorb oxygen so rapidly that the rise of temperature may destroy the wooden grids in the purifier. Some prefer to use a mixture of old and new ferric oxide so as to lessen the risk of ignition.

purifying and distributing large volumes of material; and the very valuable purposes to which the waste-products of the manufacture have been applied makes the manufacture of coal gas a process of the greatest interest. Early in the 19th century, the utilization of the large volumes of combustible gas from blast furnaces smelting iron was taken up. The uses of this gas outran the supply, and therefore C. Bischof (1839) conceived the idea of using a blast furnace solely for making the combustible by-product obtained normally from the iron furnaces.

We have seen that carbon can unite directly with two different proportions of oxygen, forming carbon monoxide and carbon dioxide. The former, carbon monoxide, can be conveniently regarded as *partially burnt* carbon; and the latter, carbon dioxide, as the final product of the combustion. One pound of carbon burning to carbon monoxide will furnish 2400 cal.; and the resulting carbon monoxide will generate 5680 cal. on combustion. Thus, one pound of carbon will produce:

Compounds formed . . .	C → CO → CO ₂
Weight changes . . .	1 → 2½ → 3½ pounds
Heat changes . . .	→ 2400 → 5680 = total 8080 cal.

In furnaces designed to make "fuel gas" by the partial oxidation of coke, the products of the actual combustion of the coke pass through a

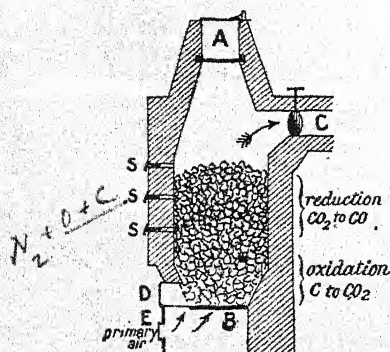


FIG. 275.—Bischof's Producer (1839).

deep bed of hot fuel. Carbon monoxide is the result. The carbon monoxide can be led to any desired spot and burnt to carbon dioxide. The furnace is called a *producer* or *generator*; and the gas coke producer gas, or coke generator gas. The solid coke in the producer is partially oxidized so as to furnish a *gaseous* fuel—hence the term *fuel gas* is sometimes used for gaseous fuels. The modern producer is a modification of Bischof's original producer in some minor details; and this, in turn, was modelled after the blast-furnace. This "persistence of type" can often be traced in many other industrial operations. In Fig. 275, A, is the charging hopper of the producer. The hopper is filled with fuel, the upper lid placed in

position; the lower shaft is drawn to the side so that the fuel drops into the producer without allowing the gas to escape. B represents the firebars, C the exit flue for the passage of the products of the partial combustion of the carbon; D is a door for cleaning firebars, etc., when required; E regulates the amount of—so-called—primary air admitted to the firebars. S represents "spy holes" for poking, etc. The shape of the producer shows that it has been modelled after the blast furnace where a combustible gas is obtained as a by-product in the smelting of iron. Instead of depending upon the blast furnace for gaseous fuel, Bischof apparently conceived the idea of making a similar furnace to supply nothing but gaseous fuel.

Since every volume of oxygen in air is accompanied by four volumes of nitrogen, coke producer gas, obviously, must contain both carbon monoxide and nitrogen. Under ideal conditions, with the volume relation $N : O = 79 : 21$, it follows that coke producer gas contains :

	Volume.	Weight.
Carbon monoxide	34.7	34.7
Nitrogen	65.3	65.3

But one pound of carbon, burning to carbon monoxide, develops 2400 units of heat. Hence 0.347 lb. of carbon monoxide, or 1 lb. of producer gas will develop 846 units of heat. This number represents the calorific power of coke producer gas. But one pound of carbon produces 6.7 lbs. of coke producer gas. Hence :

Heat from 1 lb. of carbon	8080 cals.
Heat of 6.7 lbs. coke producer gas	5665 cals.
Heat lost in conversion	2415 cals.

Hence, about 30 per cent. of the heating value of the coke is lost by the conversion of the solid coke fuel into gaseous coke producer gas. This loss is represented by the heat generated in the producer itself while

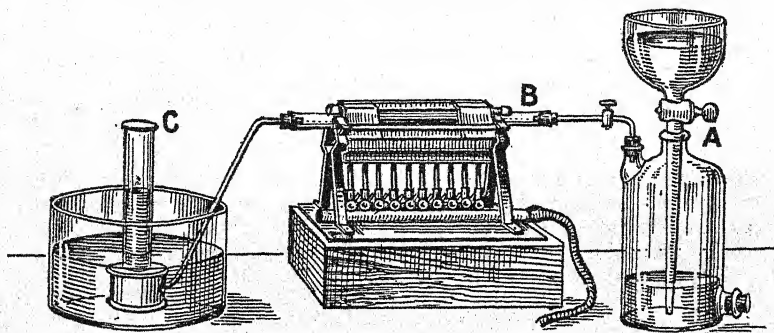


FIG. 276.—Preparation of Producer Gas.

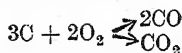
burning the coke to carbon dioxide. Industrially this loss must be counter-balanced in some way, or the use of the coke producer gas will be less efficient than direct firing with solid fuel.

Reactions in the producer.—The reactions in the producer can be imitated on a small scale in the laboratory. If a hard glass or porcelain tube *B* be packed with charcoal, and connected at one end with a gas holder containing air, *A*, Fig. 276, and the other end with a delivery tube and gas trough, *C*, when air is slowly driven through the bed of hot charcoal, carbon monoxide mixed with atmospheric nitrogen collects in the gas jar. The gas burns with a blue flame. It is coke producer gas.

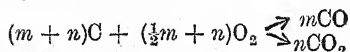
In the combustion of gaseous carbon compounds, the carbon appears to burn first to carbon monoxide (p. 921). Highly purified carbon may be heated to redness in well-dried oxygen without producing the characteristic glow of carbon in oxygen, while but a very small amount of carbon dioxide, and a large amount of carbon monoxide are obtained. We have seen,

p. 835, that dry oxygen and dry carbon monoxide unite with difficulty; and when dry, carbon burns in dry oxygen, it has therefore been inferred that carbon monoxide is the first product of the oxidation, and that the conditions of the experiment retard the subsequent oxidation of the carbon monoxide. Accordingly, when carbon dioxide is produced during the combustion of carbon in ordinary moist air, carbon monoxide is supposed to be formed first, and this immediately oxidizes to carbon dioxide: $C \rightarrow CO \rightarrow CO_2$. There is nothing to show that the mechanism of the reaction with "perfectly" dried materials as in Baker's experiment is the same as when moisture is present.

Attempts to find if carbon burns first to carbon dioxide or to carbon monoxide have not given any decisive result. In every case both gases have been detected among the products of the reaction. Hence it is not at all unlikely that we are dealing with concurrent reactions typified by decomposing potassium chlorate (p. 165):



This equation can be generalized:



by letting m denote the number of gram-atoms of carbon which burn to carbon monoxide, and n the number which burn to carbon dioxide. The relative proportion of carbon monoxide and carbon dioxide formed during the action—that is, the ratio $m:n$ —is determined by the temperature. There is also an interaction between the excess of carbon and the higher oxidation product since $CO_2 + C \rightleftharpoons 2CO$ is a balanced reaction. The relative proportion of carbon dioxide to carbon monoxide, for equilibrium, is here again determined by the temperature. For instance, any mixture of carbon monoxide and carbon dioxide when heated in the presence of carbon produces at:

Temperature.	Per cent. by volume.	
	CO	CO ₂
450°	2	98
750°	76	24
1050°	99.6	0.4

This also shows that if the temperature of a producer be in the vicinity of 450°, very little combustible gas will be obtained; and conversely, in the vicinity of 1000°, nearly the maximum possible amount of combustible carbon monoxide will be present. Hence the temperature of the producer should be about 1000° in order to get the maximum yield of carbon monoxide with a minimum loss of heat.

Coal producer gas.—If coal be used in the producer in place of coke, the coal near the top of the fuel bed will be distilled, and form coal gas. The resulting "fuel gas" will therefore be a mixture of coal gas and coke producer gas. But coal gas, as indicated on p. 865, has a relatively high calorific power, viz. 10,155 cal. Hence, the use of coal in place of coke will raise the calorific power of the producer gas. The loss in the percentage of available heat will not be so great because part of the heat is utilized in distilling the coal. One advantage of using coal producer gas

arises from the fact that slack, and inferior coal generally, can be employed under conditions where a more expensive coal would be needed for direct firing.

Burning producer gas.—The chemical engineer is constantly confronted with the "fuel problem," and in some works the composition of the producer gas and of the flue gases is regularly tested to ensure efficient working. Constructional details, type of burner, type of flame, etc., have also to be carefully studied. It is obvious that if the flue gases leave the furnace hot, and the hot gases are allowed to escape into the air, a certain percentage of heat is wasted. There are several systems for utilizing this heat to warm up the air required for the combustion of the fuel gas, etc. It will be noticed that in burning solid fuel partly in the producer, and finally in the furnace, two separate supplies of air are needed. The first supply, used for gasifying the fuel, is, for convenience, called the *primary air*, and the second supply used for burning the gaseous fuel in the furnace is conveniently styled the *secondary air*. See Fig. 274, p. 871.

The blue lambent flame which sometimes flickers over a *clear coke* (or coal) fire is burning producer gas. The air—primary air—entering at the grate reacts with the red-hot carbon: $C + O_2 = CO_2$; and this passing through the red-hot carbon of the fire is reduced: $CO_2 + C = 2CO$. The carbon monoxide on top of the fire meeting air—secondary air—burns to carbon dioxide with a blue flickering flame: $2CO + O_2 = 2CO_2$.

§ 12. Water Gas.

It will be remembered that when steam is heated in contact with metallic iron, hydrogen and iron oxide are formed; similarly, when a current of steam is blown through a bed of hot carbon (say in a gas producer), the two interact, forming hydrogen and carbon monoxide: $C + H_2O \rightleftharpoons CO + H_2$. In the latter case, it happens that *both* products of the action are combustible, and, indeed, the resulting mixture has a very high calorific power. It is called water gas. Water gas is almost free from diluting nitrogen. If the reaction occurs below 1000° , carbon dioxide begins to accumulate in the gas, and this the more the lower the temperature of the reaction. For instance, Bunte found:

TABLE LV.—EFFECT OF TEMPERATURE OF FORMATION ON THE COMPOSITION OF WATER GAS.

Temperature.	Per cent. of steam decomposed.	Percentage composition of gas produced.		
		Hydrogen.	Carbon monoxide.	Carbon dioxide.
674°	8.8	65.2	4.9	29.8
1010°	94.0	48.8	49.7	1.5
1125°	99.4	50.9	48.5	0.6

For convenience in thinking, let us suppose that the reaction occurs in two stages: (1) decomposition of the water: $2H_2O = 2H_2 + O_2$; and
2 F

(2) oxidation of the carbon by the liberated oxygen: $2C + O_2 = 2CO$. The heat required to decompose the water in the first reaction is greater than the heat given off during the combustion of the carbon by the liberated oxygen. Thus:

Heat absorbed in decomposing 18 lbs. steam . . . —58,600 cal.
Heat evolved in burning 12 lbs. of carbon to $CO + 29,520$ cal.

Heat absorbed during the reaction —29,080 cal.

Hence the producer must be getting cooler all the time the steam is passing through the fuel bed. It appears to be necessary to provide heat from an outside source to maintain the temperature of the producer sufficiently high to prevent undue amounts of carbon dioxide accumulating in the products of the reaction. It is not economical to heat the producer externally, and make the formation of water gas continuous. In modern water-gas plants, the carbon in the producer is raised to incandescence by a blast of air—the *air blow*—continued for about two minutes. This is followed by a jet of steam until the temperature falls to dull redness—*steam blow*—continued for about four minutes. When the air blow is in progress, the producer is not making water gas, and, in consequence, a damper is used to deflect the stream of gas from the producer elsewhere.

The water-gas reaction can be illustrated by substituting a flask in which water is boiling for the gas holder, A, Fig. 276. The gas which is collected can be analyzed by mixing a definite volume with air in Hempel's burette (p. 636), and exploding it in Hempel's pipette. Note the diminution in volume; and then absorb the carbon dioxide as indicated on p. 809. The data so obtained enable the amount of hydrogen and carbon monoxide to be calculated.

A rather more spectacular method of showing the reaction is to produce an electric arc in an atmosphere of steam. Water gas immediately begins to accumulate in the attached receiver.

If coke be mixed with $\frac{1}{100}$ th part of its weight of thallium chloride, and heated with water, under pressure at 350° , the reaction is represented: $C + 2H_2O = 2H_2 + CO_2$. The thallium salt acts as a catalytic agent, and the carbon dioxide can be removed by passing the mixed gases over lime. Hence, the process has been proposed as a cheap way of making hydrogen. Twice as much hydrogen is obtained by this process as in the ordinary water-gas reaction: $C + H_2O = H_2 + CO$.

A number of suggestions have been made for separating hydrogen from water gas. (1) In one proposal, the carbon monoxide is condensed by cooling; (2) another is based on the absorption of carbon monoxide by alkaline solutions under pressure when an alkali formate is produced—the hydrogen is not much absorbed; and (3) Merz and Weith suggested passing water gas over calcium hydroxide when the carbon monoxide is replaced by its own volume of hydrogen: $CO + Ca(OH)_2 = CaCO_3 + H_2$.

Semi-water gas.—By combining the operations for making producer gas and water gas—mixing the air which passes through the producer with just sufficient steam to maintain the temperature of the producer—the extra heat developed during the oxidation of carbon to carbon monoxide is utilized in decomposing the water vapour. In practice, it is found that at least 4 lbs. of carbon should be burnt by the air for every 1 lb. of carbon "burnt" by the steam. Under these conditions, the gas from a producer burning coke will be a mixture of water gas and coke producer gas. Under ideal conditions, therefore, we should have a gas containing carbon

monoxide, 37.0; hydrogen, 7.4; and nitrogen, 55.6 per cent.; and possessing a calorific power of 1144 cals. as opposed to 846 cals. with simple coke producer gas where steam is not used. Consequently, instead of losing 30 per cent. of the heat value of the fuel in the conversion, only 20 per cent. is lost.

In modern producers, the fuel gas is made by blowing steam and air into the body of the producer fed with slack coal. The result is a mixed producer gas, also called semi-water gas. For the sake of comparison, analyses of coke and coal producer gases, water gas, mixed coal producer gas, and carburetted water gas are indicated in Table LVI.

TABLE LVI.—PERCENTAGE COMPOSITION OF FUEL GASES BY VOLUME.

Constituent.		Coke producer gas.	Coal producer gas.	Water gas.	Mixed coal producer gas.	Carburetted water gas.
Non-combustible.	{ Methane	0.8 ¹	2.0	1.3	1.2	16.8
	{ Ethylene	0.0	0.4	0.0	0.2	8.7
	{ Carbon monoxide	32.3	24.4	45.6	25.2	28.7
	{ Hydrogen	4.0 ²	8.6	46.3	18.2	40.2
	{ Nitrogen	61.2	59.3	4.2	49.1	4.3
	{ Carbon dioxide	1.6	5.2	2.1	6.0	1.2
	{ Oxygen	0.1	0.1	0.1	0.1	0.1
	Calorific power (approx.)	990	1130	3560	1320	6060

The calorific powers of these gases may be compared with 10,155 cals. obtained for coal gas, pp. 866-7.

§ 13. Enriched or Carburetted Water Gas.

Water gas burns with a non-luminous flame, and, though a valuable heating agent, it is useless for lighting purposes unless it be employed in conjunction with, say, a Welsbach's mantle. Hence, if water gas is to be used as an illuminating agent, it is charged with hydrocarbon gases which do not condense on cooling. The mixture is called carburetted or enriched water gas. The carburetting is conducted as follows: When the air blast is in progress, the products of combustion from the top of the producer are deflected down a tower containing checkered brickwork, and called the *carburetter*; then up another tower also containing checkered brickwork, called the *superheater*; and thence into the air. The result of this is to raise the temperature of both towers—the carburetter and the superheater. The air valve at the top of the superheater is deflected so that the superheater is put in communication with a third tower resembling the *scrubber* of a gas works. A spray of oil is simultaneously directed into the top of the carburetter, and steam is blown into the producer. As the water gas and oil pass down the hot carburetter, the oil is decomposed—

¹ Derived from the hydrocarbons remaining in the coke.

² Derived from the moisture in the fuel and in the air.

"cracked"—and the decomposition is completed in the superheater. In this way, the oil is transformed into gases which do not liquefy when cooled. The gas is purified and washed in the scrubber, and thence passed to the gas holder.

A gas called illuminating gas, and sometimes, by courtesy, "coal gas," is a mixture of 50 to 70 per cent. of carburetted water gas with coal gas. The high percentage of carbon monoxide makes such a gas far more poisonous than coal gas. In the so-called *Pintsch's oil gas*, the oil is sprayed into hot retorts and then passed through a condenser, scrubber, and lime purifier into the gas holder.

Questions

1. What are hydrocarbons? What is a homologous series? When a hydrocarbon burns in the presence of an excess of air what are the products of combustion? How is ethylene prepared? Give the equation representing the change which takes place when ethylene burns in air. What is formed when ethylene is mixed with an equal volume of chlorine? Give the equation.—*Princeton Univ., U.S.A.*

2. How is coal gas manufactured and purified? What are the by-products? Name the diluents, the illuminants, and the impurities present in ordinary coal gas.—*Princeton Univ., U.S.A.*

3. A sample of coal contains 84 per cent. of available carbon, and 6 per cent. of available hydrogen. What weight of atmospheric air will be required to burn 1 cwt. of the coal?—*Coll. of Preceptors.*

4. How may it be proved that any given volume of ethylene (olefiant gas) contains twice as much carbon as an equal volume of marsh gas but the same amount of hydrogen?—*London Univ.*

5. 100 c.c. of a mixture of CO and CS₂ vapour were mixed with 300 c.c. of oxygen and fired. After cooling, the resulting gases occupied 340 c.c. and after absorption by potash, 200 c.c. of oxygen remained. Show how the composition of the mixture may be determined by each of the following data: (i) the contraction, (ii) the absorption, and (iii) the oxygen consumed.—*Victoria Univ., Manchester.*

6. About what proportion of the total heat given out on the complete combustion of carbon is sacrificed by first converting the carbon into carbon monoxide? Give any explanation you can of the probable causes of this difference. What do you understand by the expressions "endothermic" and "exothermic" compounds?—*London Univ.*

7. Explain as fully as you can the statement that "ethylene dibromide may be regarded either as an additive compound of ethylene, or as a substitution derivative of ethane."—*London Univ.*

8. What is meant by "saturated" and "unsaturated" compounds? Illustrate your answer by taking as examples carbon monoxide, carbon dioxide, marsh gas, and ethylene.—*London Univ.*

9. What is meant by the expression—"a homologous series"? Give an instance of such a series, with names and formulae.—*London Univ.*

10. Calculate the heat of formation of methane, CH₄, given:—C + O₂ = CO₂ + 96.9 cal.; H₂ + O = H₂O + 68.4 cal.; and CH₄ + 4O = CO₂ + 2H₂O + 213.5 cal.—*French Coll.*

11. 10 c.c. of a gaseous hydrocarbon are exploded with an excess of oxygen. A contraction of 15 c.c. is observed. After the explosion, a further contraction of 20 c.c. is observed on treating the resulting gases with potassium hydroxide solution. What is the molecular formula of the hydrocarbon?—*Customs and Excise.*

12. 18 c.c. of a gas, when mixed with 18 c.c. of oxygen, and exploded, contracted to 15 c.c. On adding potash, a further contraction of 6 c.c. took place, and the residual gas was entirely absorbed on the addition of pyrogallol. The vapour density of the gas was found to be 5.33, and none of it was capable of absorption by potash before the explosion. Draw any conclusions you can as to the

nature of the gas, and state by what experiments you would seek to confirm them.—*Oxford Univ.*

13. One volume of acetone will dissolve 25 volumes of acetylene. Assuming the validity of Henry's law, what weight of calcium carbide will be required to generate sufficient acetylene to saturate one litre of acetone at 15° , under a pressure of 9 atmospheres?—*Cleveland Univ. School, Ohio.*

14. Define acid, explaining any chemical terms you use in your definition. Discuss your definition in reference to H_2ZnO_2 , $HCCL_3$, $HKSO_4$.—*Trinity Coll.*

15. Explain the efficiency of gunpowder as an explosive. (Equation not required.)—*Massachusetts Inst. Technology, U.S.A.*

16. Define the terms: Empirical formula, structural formula, metamer, polymer, hydrolysis and homologous series, giving an example in each case.—*Owens Coll.*

CHAPTER XXXVII

ALLOTROPIC FORMS OF CARBON

§ 1. Amorphous Carbon—Lampblack.

It remains to discuss the properties of the allotropic forms of carbon.

1. The term "amorphous carbon" is used to include the different varieties of vegetable and animal charcoals—lampblack, charcoal, soot, gas carbon (p. 873), and coal. These are non-crystalline more or less impure forms of carbon. The term "amorphous," however, is rather carelessly used. Strictly speaking, the word is synonymous with "non-crystalline," but it is sometimes used in reference to the mere external irregular shape of the granules rather than to the internal crystalline structure (p. 211). It must be understood in that sense here because x-radiograms have shown that much of the so-called amorphous carbon is really crystalline like graphite.

2. Graphite includes the so-called amorphous and crystalline graphite.

3. Diamond includes boart and carbonado.

These bodies have been known for a very long time, and their appearance is so very different, that it was not until the beginning of the 18th century that chemists were reluctantly compelled to recognize that they are not only chemically related, but that they are one and all identically the same element—carbon. That these are different forms of the one element is proved by the experiment indicated on p. 806. Pure varieties of each form—sugar charcoal, graphite, and diamond—furnish on combustion the same amount of carbon dioxide per gram of material, although the heat evolved or energy degraded during the combustion of twelve kilograms of each form is different. Other characters also vary, *e.g.* :

	Diamond.	Graphite.	Charcoal.
Heat of combustion (Cals.)	94.31	94.81	97.65
Specific gravity	3.5	2.5	1.6
Specific heat	0.1469	0.2017	0.2415
Ignition temperature (oxygen)	800°–875°	650°–770°	300°–500°
Evolution CO ₂ begins	720°	570°	200°

These are therefore different forms of one element associated with different amounts of available energy.

A comparison of the atomic weights of carbon, nitrogen, and oxygen, and the large number of volatile compounds formed by carbon might make it probable that its molecule is C₂. The high boiling and freezing points, and the chemical inertness of carbon, on the other hand, make it appear as if the molecule is complex, and a comparison of the results in the above table has led to the assumption that the molecule of carbon in the diamond is more complex than in graphite, and in graphite more complex than in charcoal.

Lampblack is made by burning substances rich in carbon in a limited

supply of air so that the maximum amount of smoke is developed—for example, turpentine, petroleum, tar, acetylene, etc. The smoke is passed into large chambers in which coarse “blankets” are suspended. The “soot” collects on the blankets. Lampblack is also made from natural gas. A ring of burners is mounted below a cast iron disc with a groove on the rim convex downwards in such a way that the flame from each burner is divided into two parts. Cold water runs into the upper side of the groove, and away *via* the hollow shaft which rotates the iron disc. This keeps the metal in contact with the burning coal gas. Soot is deposited on the groove. As the disc revolves, an automatic scraper removes the lampblack from the grooves of the disc. The lampblack falls into a hopper and is conveyed by elaborate machinery to be ground to the finest powder, sifted, and weighed into sacks. Lampblack is used for making printer's ink, stove and shoe polish, paints, and in fact nearly everything in which a black pigment for colouring matter is required. Lampblack is one of the purest varieties of amorphous carbon. The analysis of a sample of acetylene “soot” furnishes 1.4 per cent. of hydrogen and 98.6 per cent. of carbon. The hydrocarbons can be removed by heating the substance in a current of chlorine.

§ 2. Charcoal.

Wood charcoal.—There are two main varieties of charcoal—wood and bone. Wood charcoal is made by burning wood with a limited supply of air in a charcoal pit or kiln; or by heating wood in closed vessels so that air is excluded. In illustration, place a few bits of wood at the bottom of a porcelain crucible. Cover the wood with a layer of fine sand so as to cut off the supply of air. Heat the crucible until combustible gases cease to be evolved. When cold, a small piece of charcoal remains in the bottom of the crucible. Note the shrinkage in volume during the carbonization by comparing a piece of charcoal with a bit of wood like that heated in the crucible preserved as duplicate. Charcoal resists the action of moisture, etc., better than wood, and hence wooden piles, fence posts, and telegraph poles are often superficially charred before being put in the ground. Some claim this treatment gives the timber a longer useful life. Charcoal is used as a fuel; in the manufacture of iron and steel; in the manufacture of gunpowder; in metallurgical operations as deodorizer; filtering medium, etc. (see below).

Pit charcoal.—In outline the industrial preparation is as follows: Small logs or billets of wood are loosely piled into vertical heaps and covered with sods and turf to prevent the free access of air. There are several systems of piling the sticks—one is illustrated in Fig. 277. A “shaft” is left in the middle of the pile to act as a central chimney or flue; and smaller holes are left round the bottom to admit the air. The pile so prepared is called a “charcoal pit” or a *Meiler* (German). The arrangement is not unlike the sulphur calcarone, Fig. 161. The wood is lighted by brushwood at the centre, and just sufficient air to allow the wood to smoulder is passed through the pile. The volatile matter escapes, and in about fifteen days the fire dies out. Between 80 and 90 per cent. of the weight of the wood, on the average, is lost by combustion, and the remaining 10 to 20 per cent. is wood charcoal. The process can only be used

satisfactorily where wood is cheap and abundant, because the method is uncertain and wasteful. The process is still employed in a few places in Europe. In Sweden rectangular piles are used, and the wood is placed horizontally and transversely.

Kiln and retort charcoal.—Some valuable gaseous and liquid products are lost in making pit charcoal. In modern processes, the wood is heated in ovens, kilns, or retorts, sealed from the outside air. The operation may be conducted simply for charcoal without recovering the by-products, or the operation may be conducted somewhat similarly to the process used for the manufacture of coal gas. The products of the dry distillation of wood include: solid charcoal in the retort; liquids—wood tar (*Stockholm tar* from pine wood); water containing wood spirit; *pyroligneous (acetic) acid*; acetone and fatty oils; and gaseous wood gas—containing hydrogen, carbon dioxide, carbon monoxide, methane, acetylene, etc. The wood



FIG. 277.—Sketch of Meiler in course of Construction.

gas is used for illuminating purposes only when the temperature of distillation has been very high. The products are approximately: charcoal, 25.3; methyl alcohol, 0.8; acetic acid, 1.0; tar, 4.0; water, 45.9; wood gas, 23 per cent. The charcoal in the retort retains the form of the wood from which it was prepared. Kiln charcoal is more compact than the pit charcoal—in the former case the charcoal weighs 20 lbs. per bushel, and in the latter case, 16 lbs. per bushel. The yield of charcoal is about 81 per cent. by volume, 28 per cent. by weight.

The dry distillation of wood can be well illustrated by placing some pieces of pine wood in a hard glass retort fitted with a receiver, etc., as shown in Fig. 278. The tar and aqueous products condense in the receiver, and the wood gas itself can be lighted. The watery liquid obtained by the dry distillation of wood is redistilled. The first portion of the distillate is the so-called "wood spirit." The wood spirit is purified by distillation from recently ignited quicklime; and by the evaporating of a mixture of the wood spirit with fused calcium chloride to dryness. The resulting compound ($\text{CaCl}_2 \cdot 4\text{CH}_3\text{OH}$) is decomposed by treatment with

water, and the solution is distilled. Finally, the distillate is rectified by repeated distillation over fresh quicklime. The resulting methyl alcohol, CH_3OH , boils at 66.78° .

Bone or animal charcoal.—This is made by heating bones, blood, etc., in closed retorts. The bones may or may not have been subjected to a preliminary extraction with naphtha or benzene to remove the fat—degreased bones; or with superheated steam or water to remove gelatine (glue)—degelatinized bones. The products of the distillation include: solid bone charcoal in the retort; liquids—a number of ammonium salts, bone oil, bone pitch, pyridine, etc.; and gases of various kinds. Bone charcoal contains about 10 per cent. of carbon, so that it is questionable if it ought to be included with the varieties of carbon at all. However, the carbon is very finely divided and disseminated through a porous mass of about 80 per cent. of calcium and magnesium phosphates, and it seems to have specially valuable qualities. Bones furnish *boneblack*—sometimes called *ivory black*—the term *ivory black* is usually applied to the product obtained by digesting bone black with hydrochloric acid to remove the calcium phosphates. Blood furnishes *blood charcoal*. For the uses of bone black and animal charcoal, see below; ivory black is used as a pigment; in the manufacture of blacking, etc. The so-called *activated charcoal*, or *active charcoal* is a form of charcoal (usually wood charcoal) prepared at a low temperature, 850° – 900° , so as to avoid graphitization. The product is specially porous and is used as an absorbent for noxious gases in gas-masks, etc.

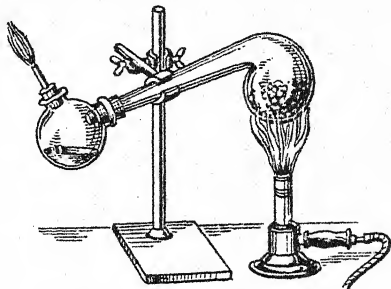


FIG. 278.—Dry Distillation of Wood.

§ 3. The Properties of Amorphous Carbon.

The specific gravity of carbon is greatly influenced by the temperature to which it has been heated, amorphous carbon varying from a specific gravity 1.45 to 1.70. Although charcoal *per se* has a greater specific gravity than water, ordinary charcoal will float on water because it is buoyed up by the air in its pores. If charcoal be weighted with a bit of lead and boiled in water for a few minutes, the air will be displaced by water, and the wet charcoal will sink when placed in cold water. A stick of charcoal can also be “anchored” below the surface of water in a tall cylinder by means of a piece of string and a weight at the bottom of the cylinder. The stoppered cylinder is then connected with an air pump, Fig. 279. The bubbles of gas rise through the water. As the air is removed, the charcoal gradually sinks to the bottom of the cylinder.

Absorption of gases.—Charcoal has a remarkable power of absorbing gases, etc. A fragment of charcoal, recently heated to expel air from its pores, is placed under a cylinder of ammonia gas, Fig. 280. The ascent of the mercury in the cylinder is a striking demonstration of the absorption of gas by the charcoal. The phenomenon is sometimes styled *adsorption*,

meaning that the gas adheres in some unknown way to the surface of the charcoal. One volume of cocoanut charcoal absorbs (Hunter), at 760 mm. pressure :

	Volumes at 0°.	Vols. at -185°.
Ammonia	171	—
Ethylene	75	—
Carbon dioxide	68	—
Carbon monoxide	21	190
Oxygen	18	230
Nitrogen	15	155
Hydrogen	17	135
Helium	2	15
Argon	12	175

At low temperatures the absorptive power of charcoal for some gases is very much augmented. Thus, a gram of charcoal which absorbs 18 c.c. of oxygen at 0°, will absorb nearly 13 times as much, namely 230 c.c. at -185°. This property affords a means of producing high vacua, and also of separating gases which are not readily absorbed (helium, neon) from those which are readily absorbed (air, etc.).

It appears as if the gases which are absorbed in greatest quantity by the charcoal are approximately those most easily condensed to the liquid state; and, rightly or wrongly, it is sometimes stated that the gases are actually liquefied on the surface of the charcoal. In any case, the "condensed" gas is usually more chemically active than the gas in the ordinary condition. Thus if charcoal be allowed to absorb chlorine, and then be brought in contact with dry hydrogen, the hydrogen and chlorine combine to form hydrogen chloride under conditions where they would not otherwise react. If charcoal which has been saturated with hydrogen sulphide be brought into oxygen gas, the rapid combination develops so much heat that the charcoal is inflamed. Sewers and foul places are sometimes temporarily purified—"sweetened"—by charcoal; charcoal biscuits have been recommended medicinally for absorbing gases in the alimentary canal in cases of flatulence.

Absorption of liquids and solids.—Charcoal also absorbs solids and liquids in a similar way. A solution of litmus (indigo, tea, vinegar, etc.) passes through filter paper without any noticeable change in the colour of the solution; but if the solution be filtered through charcoal, or if some recently ignited animal charcoal—say 10–20 grams—be shaken up with 50 c.c. of litmus solution and filtered, the filtrate is colourless. A solution of acid quinine sulphate has a bitter taste, but after filtering through animal charcoal the solution no longer tastes bitter; 10 c.c. of an aqueous



FIG. 279.—Gases absorbed by Charcoal.

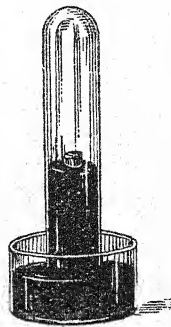


FIG. 280.—Absorption of Ammonia by Charcoal.

solution of lead nitrate (0.5 gram of the salt per litre) after boiling with 10 grams of animal charcoal and filtering, will give no precipitate with hydrogen sulphide—the original solution will. Advantage is taken of this property of animal charcoal or bone black to remove the colouring matter from many products manufactured industrially. *E.g.* coloured solutions of brown sugar are “bleached” on boiling with animal charcoal. The charcoal removes the brown resinous colouring matter, and the evaporated syrup furnishes white sugar; fusel oil can be removed from whisky by filtration through animal charcoal before the whisky is rectified. Charcoal filters are used for removing organic matter, etc., from drinking water. But since a given mass of charcoal cannot absorb an unlimited supply of organic matter, frequent cleaning is required to maintain the efficiency of the filtering medium, otherwise the charcoal charged with organic matter may serve as a culture bed for bacteria, and do harm rather than good. Consequently, the charcoal is cleansed from time to time by calcination at a red heat, otherwise it becomes clogged, contaminated, and ineffective.

Combustion.—Ordinary charcoal burns readily in air and in oxygen without smoke. The temperature at which combustion starts is largely determined by its physical condition, if the charcoal be very finely divided, it may ignite spontaneously in air. The higher the temperature to which charcoal has been heated, the higher the temperature at which it ignites in oxygen. Sugar charcoal which has been heated in the electric furnace, and graphites generally, must be heated to 660° before combustion can start between carbon and the oxygen molecules. For the vexed question whether burning carbon oxidizes first to carbon monoxide or carbon dioxide, see p. 876. One hypothesis supposes that “the first product of the oxidation is a loosely formed complex” which can be regarded as an unstable compound of carbon and oxygen of an unknown formula, and that carbon monoxide and dioxide are the primary products of decomposition of the complex at the temperature of combustion. The statement that a loosely formed complex is produced as an intermediate compound does not mean much more than saying that a certain number of carbon and oxygen molecules must come into contact before they can react.

Carbon a reducing agent.—The “affinity” of carbon for oxygen is so great that it can take the oxygen from many metallic oxides. Hence in metallurgical industries, carbon is often used as a reducing agent for ores of iron, copper, zinc, lead, etc. When a mixture of carbon with one of these oxides, say, lead, is heated in a crucible, either carbon monoxide or carbon dioxide is evolved, and the metal remains behind: $\text{PbO} + \text{C} \rightarrow \text{Pb} + \text{CO}$; or, $2\text{PbO} + \text{C} = 2\text{Pb} + \text{CO}_2$. It is interesting to compare the reaction with the action of red-hot carbon on hydrogen oxide. The water is reduced to hydrogen: $\text{H}_2\text{O} + \text{C} = \text{CO} + \text{H}_2$. It is a mere accident that both products are here gaseous.

Carbides.—The compounds of carbon with other elements are bewildering in number and complexity. Carbon unites directly with nearly all the elements to form one or more binary compounds—*e.g.* with oxygen to form oxides; with sulphur to form carbon disulphide (p. 843); with nitrogen to form cyanogen (p. 939); with hydrogen to form acetylene (p. 855); with silicon to form carborundum; and with metals to form carbides. The chemistry of the carbides may be regarded as having

been established by the work of H. Moissan. *Carbides are compounds of carbon with other elements—chiefly metals.* The most important of these, commercially, are silicon carbide and calcium carbide. Many of the carbides react directly with water forming hydrocarbons—*e.g.* beryllium and aluminium carbides give methane (p. 848); manganese carbide gives a mixture of hydrogen and methane; the carbides of the alkalis and alkaline earths give acetylene (p. 856); and the carbides of the rare earths give a complex mixture of solid, liquid, and gaseous hydrocarbons.

Calcium carbide— CaC_2 —is made by heating a powdered mixture of, say, 56 parts by weight of quicklime and 36 parts of coke in an electric

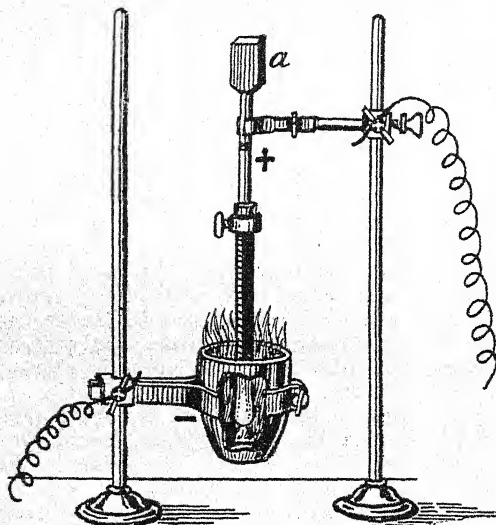


FIG. 281.—Preparation of Calcium Carbide.

arc furnace—estimated temperature 3000° —arranged so that the molten carbide can be tapped while still fluid. The endothermic action is represented by the equation : $3\text{C} + \text{CaO} \rightarrow \text{CaC}_2 + \text{CO}$. The process can be imitated on a small scale by clamping a graphite crucible to an iron rod, and connecting it with the — pole of a current of intensity 60 to 100 volts. The + pole is an electric light carbon rod which is clamped to a retort stand—Fig. 281. The retort

stands rest on some insulating material. The carbon rod is allowed to touch the bottom of the crucible and withdrawn, by the insulated handle *a*, so as to form an arc. The mixture of coke and quicklime is then gradually added to the crucible. Calcium carbide is a hard, brittle, crystalline solid, specific gravity 2.2. When pure, it is white, but commercial calcium carbide is dark grey or bronze coloured owing to the presence of impurities. Calcium carbide reacts with water forming acetylene (p. 856), and it is sold packed in tin cans to protect it from deterioration by exposure to the moisture of the atmosphere. Calcium carbide is also used in the manufacture of calcium cyanamide used as a fertilizer, and in the manufacture of cyanides.

§ 4. Coal and its Relations.

Modern civilization is the daughter of coal, for coal is to-day the greatest source of energy and wealth.—G. GAMGIAN (1913).

Vegetable origin of coal.—Geologists have potent reasons for believing that coal is of vegetable origin. The softer varieties of coal are often

changed so little that their vegetable origin is easily seen. Fossil plants can be recognized, and photographs of thin slices under the microscope show clearly the vegetable character of the coal. In some of the harder varieties, the vegetable origin can only be demonstrated by analogy and comparison with varieties less modified. There is a closely graded series ranging from peat at one end, and the anthracitic coals, or maybe graphite, at the other. It is convenient, however, to pick out certain members of the series as types. We thus obtain peat, lignite, bituminous, and anthracitic coals. There are no hard-and-fast lines between these different types; the one merges into the other by insensible gradations. The chemistry of the process appears to be somewhat as follows:

Metamorphosis of vegetable tissue into coal.—When vegetable tissue is exposed to the air, it oxidizes and decays comparatively quickly; the gaseous products of the oxidation diffuse into the atmosphere; and the mineral constituents remain behind. If the oxidation takes place in a limited supply of air, *e.g.* while submerged in a swamp or bog, the process of decomposition is rather different. Some of the

carbon is oxidized to carbon dioxide, and some of the hydrogen is oxidized to water, and probably some is transformed into methane—marsh gas—etc. As a result an increasing proportion of carbon remains behind. The total weight

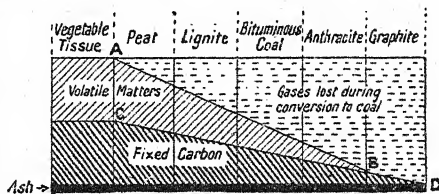


FIG. 282.—Metamorphosis of Vegetable Tissue.
(After Newberry.)

of the organic matter decreases; and, although the total amount of mineral matter—ash—remains constant, the *percentage* amount increases. These changes are represented diagrammatically in Fig. 282 (after J. S. Newberry, 1883). Assuming that vegetable tissue contains approximately 50 per cent. of carbon, and 50 per cent. volatile hydrocarbons, moisture, etc., the loss of these components, in passing to peat, is represented by the downward slopes of the lines *AB* and *CD*. These lines illustrate the changes in the proportions of volatile matters and fixed carbon in the vegetable tissue as it changes to peat → lignite → bituminous coal → anthracite → graphite. Table LVII. represents averages from between 6 and 20 published analyses of the different varieties of coals, etc., named, and they therefore represent no particular variety.

TABLE LVII.—AVERAGE COMPOSITION OF DIFFERENT TYPES OF COAL.

	Ash.	Fixed carbon.	Volatile matter.	Moisture.
Wood	1.5	25.0	53.5	20.0
Peat	1.2	29.2	51.5	18.1
Lignite	8.0	43.1	42.7	6.2
Bituminous coal	6.3	63.5	29.2	4.0
Anthracite	5.4	86.5	6.1	2.0

Origin of the different varieties of coal.—While it is probable that the early stages of the metamorphosis are brought about by bacteria and oxidation in a limited supply of air, it is also probable that the *pressure* of the superincumbent deposits of sand, mud, etc., extending over long periods of time, are needed for the later transformations. The gases—carbon dioxide, methane, water, etc.—formed during the earlier stages of the process of conversion of vegetable tissue to coal, can escape; later, when the air is shut off, methane, etc., may be imprisoned in the coal to be released as “firedamp,” p. 848, when the pressure is relieved during the mining of the coal. In some cases, the coal appears to have been *heated under pressure*. The nature of the final product, as now mined, must depend on the character of the original deposits and on the particular conditions which prevailed at the different stages of the process of transformation. The vegetable matter may have been deposited in fresh or salt water, in lakes, lagoons, seas; in marine swamps; etc. The original vegetable tissue may have been algæ deposits in sargossa seas, peat bogs, vegetable accumulations on the soil in luxuriant forests, delta and drift deposits, etc. The pressure may have been comparatively small, extended over a long period of time, and applied comparatively early in the process of transformation; the pressure may have been very great and applied late in the process of conversion; etc. An “old” coal geologically, might be “young” chemically, and conversely. Geologists can sometimes form a good idea what has happened; in other cases, they confess complete ignorance.

The different types of coal.—Assuming that peat represents the first stage in the metamorphosis of vegetable tissue into coal, it is possible to recognize several different types of peat ranging from bog-moss to heavy black peat which is closely related to lignite or brown coal—the second stage in the process of conversion. Analysis shows that lignite contains a large amount of moisture, and although it ignites readily, and burns with a long smoky flame, its calorific power is comparatively low. Lignite generally disintegrates rapidly on exposure to the air. Bituminous coal—the third stage in the transformation—is denser than lignite, black, and comparatively brittle. It seldom disintegrates on exposure to the air like lignite. Thin sections under the microscope show traces of woody fibre, lycopodium spores, etc. It burns with a yellow flame, and has a greater heating power than lignite. Some bituminous coals when heated, soften and seem to fuse, for the coal cakes into a continuous mass—**caking** or **coking** coal. Caking coals furnish a hard compact coke. Other bituminous coals do not cohere in this way when heated—**non-caking** coals. These furnish a pulverulent coke. There are all gradations between the two sub-types. Anthracite coal has a low proportion of volatile hydrocarbons, and a greater amount of fixed carbon than the other varieties. It is hard, dense, black, and brittle; it presents no trace of vegetable structure; it ignites with difficulty and burns with a short flame with a high calorific power.

Parrot or cannel coals differ from ordinary bituminous coals, and appear to have been formed differently. Cannel coals from different localities pass by insensible gradations into bituminous shales. Cannel coals burn with a luminous smoky flame, hence the term “cannel coal” (candle); they also decrepitate with a crackling noise when heated, hence

the term "parrot coal." These coals are used almost exclusively for gas making. Cannel coals yield a relatively large quantity of highly luminous gas, leaving a residue which contains a relatively small amount of coke. They contain from 50 to 70 per cent. volatile matters, 30 to 50 per cent. of fixed carbon; $1\frac{1}{2}$ to 5 per cent. of ash; up to $1\frac{3}{4}$ per cent. of sulphur; and 3 per cent. of moisture.

§ 5. Coke.

Coke can be regarded as a light porous fuel obtained as a residue when coal is heated in a closed vessel out of contact with air. Coke generally contains between 85 and 90 per cent. of carbon. Coke is used in the manufacture of iron and steel, and in a great many metallurgical operations where its comparative freedom from sulphur and certain other impurities render it more suitable than coal. Coal may be converted into coke by heating it in closed vessels—gas, tar, and ammonia are obtained as by-products; or if coal gas is being manufactured, coke, tar, and ammonia are the by-products. The properties of coke depend upon the nature of the coal from which it is obtained, and upon the way the coal is "coked." The two main varieties are *soft coke*—porous, black, brittle, ignites with difficulty; and is used for smith's forges, etc.; *hard coke*—dark grey in colour, bright lustre, compact, metallic ring when struck, bears great pressure without crushing, used for furnace work and metallurgical operations generally. Coke may or may not be prepared under conditions where the by-products are recovered. Coke is made in many ways: (1) in heaps or meilers; (2) in gas retorts; (3) in beehive ovens; and (4) in by-product ovens. The two latter are of metallurgical importance.

The *beehive oven*—so called on account of its shape—furnishes an excellent coke, but is rather wasteful. The ovens are built in batteries back to back with from 20 to 200 ovens. Each oven cokes about 7 tons of coal, and furnishes 4 or 5 tons of coke. The air for burning the coal enters through an opening in the door, and the gases escape through the top flue fitted with a damper. The air supply is diminished from day to day. When no flame is visible, and all the interior is red hot, the openings are luted with clay, and in 24 hours (70–84 hours in all) the door is opened, and water from a hose is sprayed in the oven which is then ready for discharging.

Coking in retorts—Simon-Carves' oven.—The retorts are horizontal chambers built side by side in batteries of 22 to 50. The retorts are worked in pairs—one is discharging when the other is half "coked." The retorts are closed except for the exit left for the escape of the volatile products of distillation. The products of distillation are passed through condensers, and the gases are returned to be burnt in the horizontal flues below the retorts. There is a system by which the waste heat from the products of combustion warms up the air—secondary air—which is employed for burning the gas below the retorts. In about forty-eight hours, the coke is expelled from the retort by means of a ram, and at once quenched with water. The retort is recharged through hoppers in the roof. The yield is almost theoretical. The by-products are recovered. The coke is black, hard, compact, and without metallic lustre. There are many other forms of by-product ovens. One objection to coke as a

metallurgical product is the sulphur it contains. This varies from 0.6 to 2 per cent. When this limit is exceeded, the coke is not used in many operations. The ash varies from 6 per cent. in good samples to 18 per cent. in very bad ones. The fusibility of the ash has to be considered in selecting coke for a given operation.

§ 6. Graphite.

Prior to K. W. Scheele's work on molybdenum sulphide (p. 570) when he clearly distinguished between graphite and molybdenum sulphide, the miners confused graphite with quite a host of native sulphides, molybdenum, antimony, manganese, and lead. All these substances were classed in one family because of their outward resemblance, and the fact that they all marked paper. The confusion is shown by the fact that these different substances were called molybdæna, plumbago, graphite and black-lead indiscriminately. The name molybdæna was dropped at an early date, but the other names are employed for graphite even to-day. Graphite contains no lead and the terms plumbago and black-lead are relics of the old days when all these different substances were confused together.

Graphite is widely distributed in different parts of the world. Large deposits occur in Ceylon and other parts of India, Eastern Siberia, United States, Canada, Bavaria, Bohemia, Moravia, Pinerola (Italy), etc. The mines at Borrowdale (Cumberland) are practically exhausted. Graphite also occurs in the form of fine crystals in many meteorites. The ultimate composition is represented by the following analyses:

	Volatile matter.	Carbon.	Ash.
Cingalese (Commercial) . . .	5.20	68.30	26.50
Bohemian (Schwarzbach) . . .	1.05	89.05	9.90

But more pure and less pure varieties are on the market. It was once supposed that graphite contained lead—hence graphite is sometimes called *blacklead*, and *plumbago*.

Graphite varies in specific gravity from 2 to 3. Hard graphite and soft diamonds have nearly the same specific gravity. Graphite occurs in two forms called crystalline and "amorphous." The crystalline variety has a lamellar, scaly, or flaky structure, and is largely used in the industries. The "amorphous" variety is not of much industrial importance.

Action of reagents.—Pure graphite is not attacked by heating it in a current of chlorine; nor by fusion with potassium or sodium hydroxide; some varieties are attacked by fused nitre. Chromic acid or a mixture of sulphuric acid and potassium dichromate oxidize it to carbon dioxide. Neither dilute nitric nor dilute sulphuric acid attacks graphite, although some varieties swell up into worm-like structures—sometimes 12 cm. long—when the finely granulated (not powdered) graphite is moistened with nitric acid (specific gravity 1.52–1.54) in a platinum dish, and then heated. W. Luzi (1891) calls those varieties which are indifferent to the nitric acid treatment—graphitites; and those which swell up—graphites. It is generally believed that the phenomenon is a physical effect due to the absorption of acid in the capillary pores and subsequent expansion through the development of gas under the influence of heat.

Graphitic acid.—The action of nitric acid on graphite is characteristic and distinguishes graphite from amorphous carbon, even though the

different varieties of graphite differ considerably among themselves. Finely powdered graphite is intimately mixed with 3 parts of potassium chlorate and sufficient concentrated nitric acid to give a liquid mass. After heating three or four days on a water bath, the solid residue is washed with water, and dried. The treatment with nitric acid, etc., is repeated four or five times until no further change occurs. Finally, a yellow substance is obtained which retains the form of the original graphite.¹ It is called graphitic acid (B. Brodie, 1859). The composition of graphitic acid is not quite clear.² The subject has not been investigated very much. For convenience, the above treatment is sometimes called Brodie's reaction. Diamonds are not attacked by the treatment and ordinary charcoal gives a brown mass soluble in water. Brodie's reaction is a valuable means of identifying graphite, and it is considered the safest test for graphite known at the present time. Some graphites (e.g. Bohemian) give yellow amorphous powders, others (e.g. Cingalese) give yellow microscopic lamellar crystals. If graphitic acid be heated, it swells up, forming a finely divided black powder resembling graphite, and called Brodie's graphite or pyrographitic acid. The graphite which results from the treatment of amorphous graphite is almost identical with ordinary lamp-black so far as colouring and covering power are concerned; while the graphite from crystalline graphite lacks these two qualities—colouring and covering power.

The action of heat.—Graphite when heated in air or oxygen burns to carbon dioxide, but it undergoes no change when heated in the absence of air or oxygen. It is rather difficult to ignite. The ignition temperature approaches 600°–700°. If some varieties of powdered graphite be heated in a test-tube to about 170°, the grains swell up enormously in bulk and fill the test-tube with a light amorphous powder—hence the term *sprouting graphite*.

Artificial graphite.—Iron dissolves considerable quantities of carbon, particularly if much silica be present. The higher the temperature, the greater the amount dissolved. On cooling, part of the carbon is rejected chiefly in the form of graphite (see “Iron”). Black scales of graphite can be seen on a freshly fractured surface of cast iron; and masses of it accumulate near the base of blast furnaces where it is called “kish.” Graphite is also formed when coke or charcoal is heated to a very high temperature in the electric furnace out of contact with air. Acheson's graphite is made at Niagara Falls as a secondary product in manufacturing silicon carbide, by first grinding petroleum coke with coal tar or molasses and silica or iron oxide, moulding the paste into any required shape, and baking the mixture in suitable ovens. The carbons so prepared are used for batteries, electric arc lights, etc. They are graphitized, if required, by heating in the electric furnace, when the silica and iron oxide appear to act as a catalytic agent in converting the carbon into graphite.

Uses.—Graphite when rubbed on paper leaves a black mark—hence the term graphite—from *γράφειν* (graphein), to write. It is therefore used for making *lead pencils*. For this purpose the natural graphite is purified by grinding and washing so as to remove the grit. The purified graphite is mixed with a little washed clay and forced by hydraulic pressure

¹ One treatment suffices for some graphites; others are more resistant.

² W. Luzi gives $C_{24}H_3O_{13}$; M. Berthelot, $C_{23}H_{10}O_{15}$.

through dies of the necessary shape. It is then stoved and cased in wood (red cedar for preference). Sealy graphite has been largely used, on account of its refractory qualities and high heat conductivity for the manufacture of *plumbago crucibles*. The graphite is mixed with different proportions of clay and sand—e.g. 75 parts of plastic clay, 25 sand, and 100 of graphite. The crucibles are moulded by machinery or by hand, dried, and baked at a red heat. Other refractory goods are also made from graphite. Graphite is also used as a lubricant for machinery, a coating for iron to prevent rusting, coating for goods—say plaster of Paris—to be later electrotyped, preventative for boiler scale, stove polish, polishing powder, for gunpowder, etc. Graphite is also used largely in making electric furnaces either alone or mixed with carborundum—thus *kryptol* is a mixture of graphite, carborundum, and clay. The resistance offered by this material to the passage of the electric current raises the temperature of the mass. If the mixture be suitably enclosed very little graphite is lost by combustion. Graphite conducts electricity very well, and electrodes of graphite are used in the electrochemical industries—e.g. as anodes in the manufacture of chlorine by electrolysis of sodium chloride. Graphite is also used for battery plates, electric light carbons, etc. *Colloidal graphite* obtained by deflocculating graphite with an aqueous solution of tannin and also with oil, is used as a lubricant, etc., under the commercial names *aquadag* and *oildag*.

§ 7. The Diamond.

For long ages diamonds have been prized as ornaments on account of their beauty, rarity, and permanence. As a gem, the diamond is altogether unique, for it is separated from all other gem-stones by peculiar and distinct properties—physical and chemical. The Greeks' *adamas*, the unconquerable, because it was considered to have passed unscathed through the hottest furnaces to which it had been exposed, and to resist scratching by the hardest things known. The diamond is mentioned in Exodus (xxviii. 18) as one of the 12 precious stones on which the names of the children of Israel were engraved. The ideas of the ancients about the diamond were confused and vague, and many fabulous properties were attributed to it.

Diamonds occur in their natural state as more or less rounded rough-looking pebbles not unlike pieces of gum arabic in appearance. The natural diamond must be cut and polished to bring out its lustre and sparkle. The diamond-cutting industry is developed chiefly in Amsterdam and Antwerp. The shape of the crystal as it leaves the diamond cutter has no relation to the natural crystalline shape. About half the diamond is lost during the cutting. The object of the lapidarist is to get the maximum reflection of light from the interior of the stone. The "rose" and "brilliant," for instance, are standard shapes; the latter has a rather large flat face which is really the base of a pyramid with many sides (facets). The high reflecting and refracting power of the diamond are the particular qualities which make it supreme above other gems. In virtue of these qualities, the light falling on, say, the front face of a brilliant passes into the diamond, and is reflected from the interior surface of the facets. The reflected light is refracted into a wonderful play of "lightning flashes, and sparkling scintillations" as it passes into the air. The *Cullinan*

is the largest known diamond. It was found near Pretoria (South Africa), January, 1905, and weighed over $1\frac{1}{4}$ lbs.; after cutting, it weighed about 800 carats,¹ and was valued at £100,000. It was presented to the King of England by Cape Colony in 1909. The *Regent* or *Pitt diamond* owned by the Treasury of the French Crown is the purest and finest known. It weighs $136\frac{1}{2}$ carats, and is valued at F.480,000. Stones over an ounce in weight are comparatively rare.

Occurrence.—Diamonds are sparsely distributed in different parts of the world. The chief localities are South Africa, Brazil, Ural, India, Borneo and Australia. They have been found in meteorites—*e.g.* the Canyon Diablo meteorite (Arizona, U.S.A.) contained both black and transparent diamonds. Diamonds occur in river beds and in beds or pipes containing a heterogeneous mixture of fragments of various rocks cemented together with a bluish indurated clay known as “blue earth.” The diamonds are found embedded in the blue clay. The clay crumbles on weathering, and the diamonds are readily detected in the disintegrated mass.

Varieties.—Diamonds are usually tinged slightly yellow. The clearest and most nearly colourless diamonds without flaw are most prized as “diamonds of the first water.” Diamonds are also occasionally coloured blue, pink, red, and green owing to the presence of traces of foreign metals. Some diamonds are dark grey and even black; they exhibit a more or less imperfect crystalline structure, and are known as *black diamonds*—*boart* or *bort*, and *carbonado*. *Boart* is an imperfectly crystallized black diamond which has various colours, but no clear portions, and is therefore useless as a gem; *boart* is used in the drilling of rocks, and in cutting and polishing other stones. *Carbonado* is the Brazilian term for a still less perfectly crystallized black diamond. It is as hard as *boart*, and has similar uses. *Boart* and *carbonado* are usually regarded as intermediate forms between diamonds and graphite.

Properties.—The diamond is rather brittle. It is the hardest substance known. Crystalline boron comes next; it is nearly as hard as the diamond. The hardness, refracting power, and other properties vary with different diamonds; and, indeed, in different parts of one diamond. The specific gravity varies from 3.514 to 3.518; *carbonado*, 3.50; *boart*, 3.47 to 3.49. Amorphous graphite has a specific gravity of 2.5; hard gas carbon, 2.356; and amorphous carbon, 1.45 to 1.70. The diamond is transparent to Röntgen's rays, whereas glass, used in imitation of the diamond, is nearly opaque to these rays. This furnishes a ready means of distinguishing imitation diamonds from the true gems.

The diamond is insoluble in all liquids. Fused potassium hydrofluoride mixed with 5 per cent. of nitre attacks the diamond slightly; a mixture of potassium dichromate and sulphuric acid oxidizes the diamond to carbon

* Diamonds are sold by the “carat.” The carat represents the Roman *siliqua* or Greek *κεράριον*. It was $\frac{1}{4}$ th of the golden solidus of Constantine, which was $\frac{1}{4}$ th oz. It is now a measure of weight used for diamonds and other precious stones, and was originally $\frac{1}{144}$ th oz. or $3\frac{1}{2}$ grains, but it is now nearer 0.207 gram, $\frac{1}{160}$ th oz. troy, or $3\frac{1}{2}$ grains troy. The value of the international carat is 0.2 grm., or 3.08647 grains troy. The term “carat” is derived from the carob bean, formerly used as a small weight by the diamond merchants of India. Its exact value has varied in time, and it varies in different countries. The carat is also a proportional measure of $\frac{1}{24}$ th, and is used in stating the fineness of gold.

dioxide at about 200°. Unlike graphite and amorphous carbon, diamonds are scarcely attacked by a mixture of potassium chlorate and nitric acid.

The action of heat.—If a clear crystal of the diamond be placed between two carbon poles of an electric arc—in the absence of air—when the temperature approaches 2000°, the diamond swells up and changes to black graphite: $C_{\text{diamond}} = C_{\text{graphite}} - 0.5 \text{ Cal}$. The diamond, boart, and carbonado commence to burn when heated in air between 710° and 900°. The ignition temperature depends upon the hardness, etc., of the diamond. A soft Brazilian diamond burnt at 710°, and an exceptionally hard boart at 900°. It is rather difficult to burn the diamond



FIG. 283.—The Combustion of Diamonds in Oxygen.

unless the temperature be maintained by, say, placing the diamond on a piece of platinum foil heated red hot by an electric current. An apparatus suitable for burning the diamond in oxygen is illustrated in Fig. 283. A splinter of diamond is placed on a piece of stiff platinum foil arranged so that it can be heated red hot by the passage of an electric current when the globe is full of the required gas. Some varieties of graphite require a special method of ignition before they can be burnt in oxygen gas. The product of the combustion is carbon dioxide (p. 800). A light ash consisting of iron, lime, magnesia, silica, and titanium remains. Clean crystalline diamonds have about 0.05 per

cent. of ash, whereas with boart the ash may run as high as 4 per cent. Arguing from the high refractory power of camphor, olive oil, amber, etc., "which are fat, sulphureous, unctuous bodies," Isaac Newton (1675) inferred that "a diamond is probably an unctuous substance coagulated," although he was anticipated by A. Boetius de Boot in 1609; thus predicting what was later on proved experimentally. In 1694, Averani and Targioni of the Academy of Amento first burnt the diamond by exposing it in the focus of a large "sun glass." Lavoisier in 1775, and H. Davy

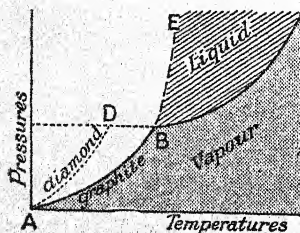


FIG. 284.

in 1814, showed that the diamond forms carbon dioxide when burnt, and unlike graphite and charcoal, it gives no aqueous vapour, showing that it is free from hydrogen and water. Hence it was inferred that the diamond is crystallized carbon.

The synthesis of diamonds.—Molten solids, on cooling, generally crystallize. Carbon, however, volatilizes at ordinary atmospheric pressures at about 3600° without passing through an intermediate liquid state. Arsenic also

volatilizes at ordinary pressure without liquefying; but arsenic easily liquefies if it be heated under pressure. It is therefore inferred that if sufficient pressure could be obtained, carbon also would melt to a liquid which would crystallize on cooling. The diamond at ordinary temperatures is supposed to be the unstable, and graphite the stable form. If the one form does pass into the other at ordinary temperatures, the speed of transformation must be extremely slow. A. Smits' hypothetical views (1905) are summarized in Fig. 284. It is estimated that if the pressure

were great enough, graphite would melt at about 3000° , and that the liquid would have a vapour pressure represented by BC , Fig. 284, and the solid a vapour pressure AB . The vapour pressure of the diamond would be represented by AD where D is the melting point of the diamond under great pressure. BE shows how the melting point of carbon changes with pressure. If molten graphite be under-cooled the solution would become more and more viscous, and appear as amorphous carbon.

Iron dissolves carbon and gives it up again on cooling. Other metals, especially silver, behave in a similar manner, but iron appears to be the best solvent. The solubility increases with the temperature. Moissan packed a piece of iron, as pure as practicable, in a carbon crucible with sugar charcoal. The crucible was heated between the poles of an electric arc furnace (700 amps., 40 volts)—Fig. 285. Under these conditions the iron melted and dissolved much carbon. When the temperature had reached 4000° , and the iron was volatilizing in clouds, Moissan plunged the crucible in cold water. The sudden cooling solidified the outer layer of iron. The expansion which the inner liquid core underwent on solidifying must have produced an enormous pressure. Hence, the carbon separated from the iron under a very great pressure. After dissolving away the iron, etc., some of the carbon which remained was in the form of boart—black diamonds—some as graphite, and some in the form of transparent diamonds—microscopic, it is true. Moissan separated as many as 10 to 15 minute transparent diamonds from a single ingot treated in this way. The largest was about $\frac{3}{4}$ mm. long. Crookes also appears to have detected diamonds in the carbonaceous residue obtained when cordite is exploded in closed steel cylinders where the pressure is estimated to be as high as 8000 atmospheres and the temperature over 4000° .

E. de Boismenu (1907) claims to have synthesized diamonds approaching 2 mm. in diameter by the electrolysis of molten calcium carbide. The diamonds were found about the anode. A kilogram of the carbide is said to have furnished 1.75 carats of true diamonds. The claim is unconfirmed.

§ 8. Electric Furnaces.

The use of electrical energy for metallurgical furnaces is becoming of greater importance every day. The heat required for the reduction of the metals from their ores, and for the subsequent treatment of the metals is no longer exclusively produced by the combustion of carbonaceous fuels—charcoal, coal, coke, gas, petroleum, etc. Carbonaceous fuel is apparently used in some electric smelting furnaces, but rather as a reducing agent and carbonizer than as a source of heat. The most important characteristic of electric energy is the concentration of almost any degree of heat in a given space so that operations can be readily performed which are impossible at the lower temperatures available with carbonaceous fuels. The applicability of the electric furnace for ordinary metallurgical operations—reduction of iron ores, etc.—is feasible where cheap water power is available and fuel is costly, although usually, electrical power is too costly for ordinary smelting operations.

It seems here advisable to emphasize the difference in some types of electric furnace indicated in this work. Furnaces in which the electric current is used as a source of heat are not to be confused with those in which

the electric current is used for electrolysis. The main types of electric furnace are :

1. Arc furnace.—In this, the heat is produced by one or more electric arcs. The arc may be established between one or more pairs of carbon or graphite poles as in Moissan's furnace—Fig. 285 ; or between the fused metal bath or fused slag and a carbon pole as in some furnaces used for reducing iron ores to pig iron, and in refining pig iron and steel.

2. Resistance furnace.—Here the heat is produced by the passage of an electric current through a solid or liquid resistor.

(a) A special resistance, *e.g.* nickel, ni-chrome, platinum wire, fragments of carbon, etc., is embedded between the inner and outer walls of the

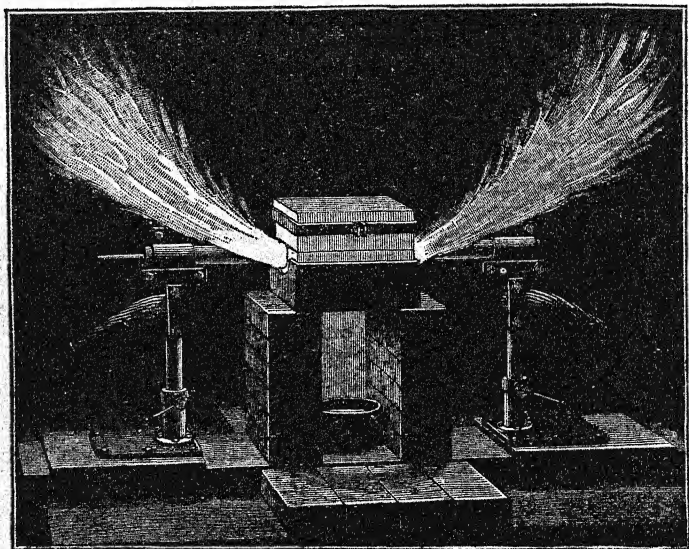


FIG. 285.—Moissan's Electric Arc Furnace.

furnace. The inner wall may take the form of a muffle, tube, etc. These furnaces are fairly common in chemical laboratories.

(b) The charge in the furnace constitutes the resisting medium. The resistance of the medium raises the temperature of the charge, *e.g.* the phosphorus furnace (Fig. 234); the calcium carbide furnace (Fig. 281); the carbon disulphide furnace (Fig. 264); and the carborundum furnace (Fig. 331).

3. Electrolytic furnace.—A continuous current splits the fused electrolyte into its component parts. The heating effect of the current may or may not suffice to keep the contents of the furnace in a fluid condition. The aluminium furnace (Fig. 246) illustrates the former; and Castner's sodium furnace (Fig. 147) the latter.

4. Transformer or induction furnace.—The molten metal forms part of the secondary of a large transformer, and the current is not led in by terminals or electrodes, but by wireless transmission is generated directly

in the metal bath. The furnace hearth is thus connected with the dynamos in no visible way. These furnaces are used in steel refining, and are not illustrated in this text-book.

Questions.

1. What do you understand by the terms element and compound? What experiments would you make on (a) ferric oxide, and (b) carbon in order to determine in the case of each of these substances whether it is an element or a compound?—*Univ. North Wales.*

2. You are given a jar containing one of the following gases:—hydrogen, oxygen, nitrogen, nitrous oxide, carbon monoxide, methane, and carbon dioxide. Describe exactly how you would proceed to ascertain which gas is present.—*Sydney Univ.*

3. Name the principal "impurities" which are usually present in the air of towns. How could you determine the amount of any two of these substances present in 100 volumes of air? How does town air differ from country air, and how may these differences be accounted for?—*Board of Educ.*

4. Describe properties of carbon which tend to show that diamonds cannot have been formed at a temperature at which pure iron melts. How would you show that carbon dioxide is a compound of carbon and oxygen, and that it contains very nearly its own volume of oxygen?—*London Univ.*

5. Carbon is "tetravalent" in many of its compounds. Explain the meaning of this statement.—*London Univ.*

6. State what happens when the following substances are heated to a sufficiently high temperature with charcoal: sulphur, phosphoric oxide, arsenious oxide, zinc oxide, litharge, sodium carbonate. Express the reactions in equations.—*Science and Art Dept.*

7. What is charcoal, and how can it be made on a large scale? Describe the difference between peat, lignite, bituminous coal, anthracite, and graphite.—*Bombay Univ.*

8. If 100 parts by weight of air contain 23 of oxygen and 77 of nitrogen, how many tons of air would be needed to burn one ton of coal consisting of:

Carbon.	Hydrogen.	Nitrogen.	Oxygen.	Ash.
90.55	4.14	1.26	2.35	1.70

Assume that, on burning, the carbon is wholly converted into carbon dioxide, and the hydrogen into water; that the nitrogen is evolved as such, and that the ash is already fully oxidized.—*London Univ.*

CHAPTER XXXVIII

COMBUSTION AND FLAME

§ 1. Mayow's Work on Combustion.

Slowly, gradually and laboriously one thought is transformed into a different thought, as in all likelihood one animal species is gradually transformed into a new species. Many ideas arise simultaneously. They fight the battle for existence not otherwise than did the *Ichthyosaurus*, the Brahmin, and the horse. Thoughts need their own time to ripen, grow, and develop.—E. MACH.

SEVERAL of the early writers knew quite well that fire cannot exist without air, thus, in the 13th century, Roger Bacon said that "air is the food of fire"; and near the beginning of the 16th century, Leonardo da Vinci clearly recognized that air is necessary for the sustenance of the flame of a burning candle, for he said "there is smoke in the centre of the flame of a wax candle because the air which enters into the composition of the flame cannot penetrate to the middle. It stops at the surface of the flame and condenses there." Leonardo considered that air is not an element because one part of it alone is concerned in combustion.

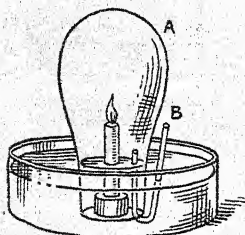


FIG. 286.—Mayow's Experiment on Combustion.

R. Boyle (1661) also noticed that owing to the "want of air" the flame of a lighted candle expired more quickly under the exhausted receiver of an air pump than when the receiver was not exhausted. This and other experiments on similar lines showed that, air is necessary for combustion. Robert Hooke (1665) suggested that air was mixed with a substance "which is like, if not identical with, that which is fixed in saltpetre."¹ Hooke also seems to have recognized that only this portion of air is concerned in combustion. John Mayow (1674) subjected the guess or hypothesis of Hooke to the test of experiment. Mayow arranged a candle in water so that the wick was between 9 and 10 cm. above the surface of the water. A glass cylinder, A, Fig. 286, was lowered over the burning candle so that the level of the water inside and outside the cylinder was the same.² The flame of the candle soon expired, and water rose in the jar. Some gas

¹ Scheele later produced oxygen from saltpetre—that is, potassium nitrate.

² Mayow describes a small syphon B, Fig. 286, which he used for this purpose. Immediately the jar was in position, the syphon was removed.

still remained in the jar, but it could not be air because one of the characteristic properties of air is to support the burning of the candle, and the flame of a candle is immediately extinguished in the residual gas. Hence Mayow inferred that air contains two kinds of particles one of which—the “nitro-aërial particles”—is “withdrawn and destroyed”¹ by the burning candle. Mayow does not seem to have quite grasped the idea that “the nitro-aërial particles” which support combustion actually combine with the burning body, although he correctly inferred that the air was a mixture of nitro-aërial and aërial particles.²

In fine, Mayow's observations show that air is a mixture of two gases, one of which is withdrawn during combustion, and the remaining gas does not support combustion. Mayow got very near to the present-day theory of combustion, but unfortunately, his ingenious experiments had very little, if any, influence on the subsequent development of chemistry, because, for another century, more trust was placed in phantasms of the imagination than in facts obtained by precise observations.

§ 2. The Phlogiston Theory.

During the greater part of the 18th century, the doctrine of phlogiston was not only the lamp and guide of chemists, but it remained the time-honoured and highest generalization of physical chemistry for over half a century.—S. P. LANGLEY.

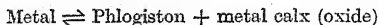
Phlogiston died as an old king—once infinitely dominant, somewhat tyrannical, not always just; now deposed, decrepit, utterly senile, forsaken by all.—W. ODLING.

The beautiful fiction of Greek mythology, as related by Æschylus, tells how Prometheus stole fire from heaven and gave the sacred gift to man as the most useful of all his necessities. To many ancient worshippers fire was a thing divine, the supreme manifestation of God Himself; and it soon became the one visible symbol of God Who was called by the Zoroastrian fire worshippers “the one fire.” Fire thus came to be the first and most potent of all the elements, and it is thus easy to understand how Heracleitus came to regard fire as the sole primal element from which all things were created; and how fire was canonized by Pythagoras and Empedocles as one of the four indispensable and all-sufficient components of the whole universe. Accordingly, about the middle of the 18th century, combustion was explained by the aid of the assumption that all combustible substances contained the essence of fire—an inflammable principle—which enabled them to burn. This obviously means very little more than that substances burn because they are combustible. Geber (c. 770) thought that the inflammable principle must be sulphur—*ubi ignis et calor, ibi sulphur* (“where there is fire and heat, there is sulphur”). J. J. Becher (1669) pointed out that many combustible substances were known which did not contain sulphur, and he was led to postulate the existence of another principle which he termed *terra pinguis*—fatty or inflammable earth. Becher's fatty earth became Stahl's phlogiston—from the Greek *φλογιστεω* (phlogisteo), I set on fire. G. E. Stahl (1723) taught that in the act of combustion phlogiston, an intrinsic constituent

¹ “... *ab aere exhauriri et absumi.*”

² Some modern commentators consider the former to be oxygen, the latter nitrogen.

of every combustible body, is set at liberty. Oxidation was said to be due to the escape of phlogiston; reduction to the absorption of phlogiston. When a metallic oxide was heated with a substance rich in phlogiston, *e.g.* charcoal—or reducing agents generally—the charcoal supplied the calx or metallic oxide with phlogiston, and reproduced a compound of phlogiston with the metallic oxide which was the metal itself. Metals were thus supposed to be compounds of phlogiston and their calces (oxides). The noble metals have their calces so firmly fixed that nothing can take it from them. While the base metals are turned into calces when roasted in air, the royal metals remain intact during the fiercest trial. If phlogiston escaped, the metallic oxide remained. The idea can be symbolized:



When it was shown that the metallic oxides were heavier than the corresponding metals (*e.g.* Rey's experiment) it was assumed that phlogiston was lighter than air, so that the metal was buoyed up, so to speak, by the associated phlogiston (p. 20). Phlogiston began as an imponderable principle, and it narrowly escaped being generally identified with hydrogen, and some did actually consider hydrogen to be phlogiston itself; one of the old names for hydrogen was phlogisticated air.

Lavoisier's conclusive proof (1774) that the increase in weight which occurs during oxidation is equal to the weight of oxygen absorbed from the air; and his crucial demonstration that combustion is a process of absorption and increase in weight which cannot be explained by a supposed loss of substance, soon banished the phlogiston hypothesis from the domain of science. Hence it has been said that chemistry was not born until burning was investigated. It is not surprising that writers on chemistry in the Middle Ages failed to interpret the experiment of the burning of a candle in air when we recall the knowledge required to explain the chemical side of the phenomenon, altogether apart from the skill required in the manipulation of gases:

- (1) Air is composed of two gases both sparingly soluble in water;
- (2) During combustion, one of the gases unites and the other does not unite with the burning body;
- (3) Air contains four volumes of the inert gas, and one volume of the gas which unites with the burning body;
- (4) A gas soluble in water is produced during the combustion; and
- (5) The increase in weight of the combustible body during combustion is equal to the decrease in the weight of the air.

The phlogiston hypothesis is sometimes held up to ridicule. We must bear in mind that the hypothesis was adopted by nearly all the leading chemists in the earlier part of the 18th century when it appeared to be as firmly fixed among the root principles of chemistry as the kinetic theory does to-day. Thus Priestley said:

If any opinion in all the modern doctrine of air be well-founded, it is certainly this, that nitrous air is highly charged with phlogiston. If I have completely ascertained anything at all relating to air, it is this.

The phlogiston theory represented the most perfect generalization known to the best intellects of its day. Under its influence chemistry thrived and multiplied its proportions. It is inconceivable that men like Bergmann, Black, Cavendish, Priestley, and Scheele would counsel what they con-

sidered to be an inconsistent doctrine. Phlogiston was regarded, not as a temporary hypothesis, but as a permanent acquisition, an enduring conquest of truth. To-day, the word is but an empty symbol.¹

Theories perish, facts remain.—Much of what we think best in the theories of to-day, may to-morrow be rejected, with phlogiston, worthless. It is not at all improbable that posterity will be amazed at some examples of our inexplicable ignorance. This need cause the student no embarrassment. A fallacious theory may be a valuable guide to experiment. Experiment and labour applied to the explication of the most extravagant hypothesis is not always lost. Guided by wrong hypotheses men have sought one thing and found another—Columbus sought the Indies, and discovered America—but it is necessary to follow René Descartes' advice : Give unqualified assent to no proposition which is not presented to the mind so clearly that there is no room for doubt. As Aristotle would have said, we do not need to cultivate the art of doubting, but rather the art of doubting well.

§ 3. Is Combustion Oxidation ?

When a theory which has been received on good evidence appears to fall, the really essential and vital part of it survives the fall ; that which has been discovered continues to be true.—W. WHEWELL.
Inflammable air is undoubtedly charged with abundance of the principle of inflammability.—T. BERGMANN (1779).

Lavoisier's work in 1774, on the composition of air, and on the increase in weight which occurs when a metal is calcined in air, has already been described in outline. During the next two years, Lavoisier proved that carbon dioxide is the product of the oxidation of carbon—diamond and wood charcoal—and that carbon dioxide and water are the products of the combustion of organic compounds containing only carbon, hydrogen, and possibly oxygen. In 1777 Lavoisier published his oxidation theory of combustion : (1) Oxygen is necessary for combustion ; (2) Oxygen is consumed during combustion, and the increase in weight of the substance burnt is equal to the decrease in weight of the atmospheric air. The origin of the water formed during the combustion of hydrogen was completely explained by Cavendish's synthesis of water in 1783. Consequently, combustion is a process of oxidation which is attended by the development of light and heat. When carbon burns in air, carbon dioxide is formed ; when hydrogen is burnt in air, water is formed ; when phosphorus burns in air, phosphorus pentoxide is formed ; sulphur furnishes sulphur dioxide ; etc. Lavoisier's work rendered the old *principle of inflammability* redundant in the same sense that Newton's work on gravitation rendered the *principle of ponderosity* unnecessary to describe the falling of bodies.

It must be added that many other chemical reactions which furnish light, heat, and flame, are now also included under the term "combustion," even though oxygen be absent. Hence combustion is not always

¹ A. L. Lavoisier (1789) considered oxygen to be made up of caloric and the matter of oxygen. H. C. St. Deville (1860), Crum Brown (1866), and W. Odling (1871) have pointed out that "phlogiston" occupied a similar position in the chemistry of the 18th century that "potential energy" does to-day. Here, then, the old revives in the new. The chemistry of to-day is not materialistic, for it is occupied with both energy and matter.

oxidation, for the term is applied generally to any chemical process which is attended by the development of heat and light. For instance, the "combustion" of brass in chlorine; of hydrogen in chlorine, etc. The development of flame, light, and heat during combustion is quite an accidental feature of the process of oxidation. The speed of the oxidation may vary from the slow decay of organic matter occupying maybe centuries, to the rapid inflammation—explosion or detonation wave—travelling at the rate of 10,000 feet per second.

Pyrophoric powders.—Finely divided lead is prepared by heating lead tartrate, at as low a temperature as possible, in a glass tube sealed at one end. When the evolution of gas has ceased, either close the open end with a cork, or seal it up hermetically while hot. If the cold powder be allowed to fall through the air on to the floor, the oxidation of the powder proceeds so rapidly that the temperature is raised and the falling powder becomes red hot.¹ Hence the term *pyrophoric lead*. Finely divided iron oxalate, nickel oxide, etc., reduced at a low temperature in a stream of hydrogen, also furnish pyrophoric powders.

Spontaneous combustion.—The so-called spontaneous ignition of coal which takes place in gob,² in coal stacks, in coal bunkers, etc., is said to be caused by the oxidation of pyrites in the coal; it may also be due to the oxidation of the coal itself,³ owing to excessive ventilation leading to a more rapid supply of air, and the oxidation of pyrites is in many cases of but minor importance. Heat is generated. In some cases, if ventilation be defective, the heat is not dissipated, but warms up the mass; the speed of oxidation is accelerated by the rise in temperature; this generates still more heat until finally the mass ignites. Some coal dust mixed with air ignites in the vicinity of 200°.

Oxidation of oils.—Many oils—linseed oil, nut oil, poppy oil, hemp oil—absorb oxygen. Linseed oil used for paint, for instance, absorbs oxygen and forms a tough resinous skin as the paint "dries." This skin holds the colouring matters in suspension and protects the material underneath. The paint oils, therefore, do not "dry," in the ordinary sense of the word, by evaporation, but by oxidation. The heat generated during the oxidation is dissipated; but if rags, etc., greasy with oil, be left in a badly ventilated place, or if a heap of greasy material be left for some time undisturbed under such conditions that the mass is not well enough ventilated to keep it cool, the temperature may rise high enough for spontaneous inflammation. The oxidation of the paint oils is facilitated by the addition of lead and manganese salts.

Eremacausis.—The decay of wood is a process of oxidation, very slow, it is true, and it proceeds by what appears to be a roundabout process. The end products, however, so far as we can tell, are the same as those obtained by direct combustion. The heat of the reaction too is no doubt the same (Hess' law) whether the wood decomposes slowly with the formation of a large variety of intermediate products, and after a number of

¹ The experiment is more amusing if the powder be allowed to fall on a *small* heap of gunpowder.

² "Gob" is the space from which coal has been removed, and where the "slack" is usually left behind as worthless.

³ There are other causes of underground fires. Thus, the excessive ventilation of the mine may lead to gob fires owing to the greater supply of oxygen, and consequently more rapid oxidation.

years ending in carbon dioxide and water; or whether it proceeds more directly in a few minutes by combustion. In both cases the same amount of energy is degraded although the process of degradation proceeds by different reactions. The process of slow oxidation without the application of heat is sometimes called **slow combustion** or **eremacausis**—from the Greek *ἡρέμα* (*erema*), quietly; *καύσις* (*causis*), to burn. The rusting of iron (*q.v.*) is another example of a slow oxidation, although the end product is not the same as in the combustion of iron in oxygen. In the former case a hydrated ferric oxide— $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$ —is formed, and in the latter case magnetic iron oxide— Fe_3O_4 . The absorption of oxygen by rusting iron is generally illustrated by inserting a plug of steel wool, *A*, Fig. 287, into a 100 c.c. eudiometer tube *B*, resting in a dish of water *C*. The level of the water *B* is adjusted, and the apparatus examined after about twenty-four hours. The iron will be coated with rust, and the enclosed air will have decreased in volume. A stick of phosphorus at the end of a wire may be held in the cylinder in place of the iron. The result of the experiment will be similar—removal of oxygen.

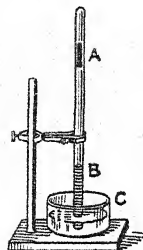


Fig. 287.—Iron Rusting.

In the "dilution method" of sewage treatment, the sewage is mixed with a large volume of water. The oxygen dissolved by the water, assisted by bacteria, quickly renders the organic matter innocuous mainly by oxidizing it into carbon dioxide and water—and possibly nitrogen compounds.

§ 4. Respiration involves Oxidation.

A living being is at work; work means waste; waste means decay; decay is combustion; combustion is oxidation.—E. AVELING (1891).

John Mayow (1674) demonstrated that the respiration of animals is a process of oxidation analogous with combustion. Mayow placed a mouse in a cage under a vessel standing over water—Fig. 288—and noticed that the water rose in the jar as respiration continued just as if a burning candle had been placed under the jar—p. 900, owing to the withdrawal of "nitroaërial particles" (oxygen). He found that the mouse died after a time, and it was impossible to ignite a combustible body in the residual "aërial" gas (nitrogen).

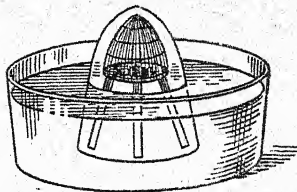


Fig. 288.—Mayow's Experiment on Respiration.

Metabolism of animals and plants.—All living organisms are continually wasting tissue as a result of muscular or other forms of work.¹

¹ Those constructive processes by which substances taken in as food by living organisms are converted into protoplasm and their own proper substance, are said to be anabolic—from the Greek *ἀνά* (*ana*), anew; *βάλλειν* (*ballein*), to throw. Those destructive processes by which the protoplasm, etc., breaks down into simpler products are termed katabolic—from the Greek *κατά* (*kata*), down. The two phenomena are included under the term metabolism—Greek *μετά* (*meta*), about.

The air during respiration loses about 4 to 5 per cent. of oxygen, and gains about 3 per cent. of carbon dioxide. Thus, the percentage composition of inhaled and exhaled air is reported by physiologists to be as follows:

	N ₂	O ₂	CO ₂	Moisture.
Inhaled air . . .	79.15	21.00	0.04	Varies.
Exhaled air. . .	79.59	16.00	4.38	Nearly saturated.

Oxygen, taken in by the lungs, is absorbed by the blood. The blood contains hæmoglobin, this unites with oxygen, forming oxyhæmoglobin—the former predominates in blue venous blood; the latter, in red arterial blood. The oxyhæmoglobin gives up its oxygen very readily, and it thus oxidizes the waste products—partly to carbon dioxide. The carbon dioxide held in solution by the venous blood is pumped by the heart to the lungs, and the gas is finally exhaled in the breath. Thus, with each respiration the animal robs the atmosphere of some of its oxygen, and introduces into it a proportionate quantity of carbon dioxide.

By breathing into a beaker, and then adding clear lime-water; or better, by blowing through a glass tube into a beaker containing clear lime-water, it is easy to demonstrate the presence of carbon dioxide in the breath. The precipitation of calcium carbonate causes the clear lime-water

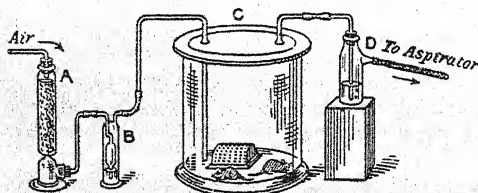


FIG. 289.—Exhalation of Carbon Dioxide by Animals.

to become turbid, etc. The experiment can be modified by drawing a stream of air through pieces of "soda-lime" in a tower, A, Fig. 289, to remove carbon dioxide from the air; then through a washbottle, B, with clear lime-water. The clarity of the lime-water demonstrates the absence of carbon dioxide in the stream of air. The air after passing the *temoin* (witness) tube, B, enters a glass chamber, C, which holds a mouse or a couple of mice, and then passes through another washbottle, D, containing clear lime-water. The turbidity of the lime-water in D, and the clarity of the lime-water in C, can only be explained by assuming that the mice are giving off carbon dioxide during respiration. The same vessel C can be used for trying the effect of different gases on mice.

Less carbon dioxide is exhaled during sleep because less waste products are formed. The heat evolved during the oxidation processes in the body maintains the body at the necessary temperature. The respiration of animals and plants is essentially a true process of combustion, for in order to secure the necessary warmth, a part of the body is burnt, and the necessary supply of oxygen and the removal of the products of combustion—carbon dioxide, etc.—is effected by the lungs in breathing. In the case of plants, most of this work is done by the leaves. The leaves of plants thus act as so many lungs. The respiration of animals and plants goes on continuously, night and day—oxygen (air) is taken in, and returned to the air as carbon dioxide and water. The consequent loss in weight in both animals and plants is made good by assimilation or feeding. Fish abstract dissolved oxygen from water by the aid of their gills. Fish quickly die for want of oxygen when placed in cold water which has been deprived

of dissolved "air" by boiling. Aërated water¹ furnishes the oxygen necessary for their sustenance.

Similar remarks apply to plants, and the abstraction of oxygen from air during the respiration of plants is illustrated by the experiment, Fig. 290.

Assimilation by plants.—The respiration of plants must not be confused with the process of assimilation. In daylight, plants absorb carbon dioxide from the air, fix the carbon, and give off oxygen. In sunlight, the process is very active, so that the respiration process, though at work, is masked because more carbon dioxide is taken in and oxygen given out by assimilation (feeding) than oxygen is taken in and carbon dioxide given out by respiration (breathing). The decomposition of the carbon dioxide (feeding) only occurs in daylight, and plant life is thus dependent upon this process for nourishment. This action of light may be readily shown by loosely packing a two-litre flask with green leaves, and filling up the flask with water through which carbon dioxide has been allowed to bubble. Fit the flask with a stopper, and funnel as shown in the diagram, Fig. 291. Also fill the flask and part of the funnel with water. A test-tube inverted over the funnel will collect sufficient oxygen to allow the usual tests being made.²

The energy required for the decomposition of the carbon dioxide is derived mainly from the solar light and heat. There is thus a transformation of the sun's energy into chemical energy which is stored up as vegetable tissue. The heat of combustion of burning wood is believed to be equivalent to the solar energy used in its formation (Hess' law). At night, in darkness, feeding stops, but breathing—the absorption of oxygen and the evolution of carbon dioxide—continues. The net result of both processes—assimilation and respiration—is to remove carbon dioxide from the air, and restore oxygen. There is probably a kind of equilibrium pressure between plants and the carbon dioxide of the air. If the amount of carbon dioxide exceeds a certain limit, this would be followed by greater activity in vegetable life, and thus the normal proportion of carbon dioxide would be restored. Similarly, if the amount of carbon dioxide were reduced below the normal, vegetable life would be impoverished until equilibrium was restored. The proportion of carbon dioxide present in the gases dissolved by sea-water is about 27 times as great as the proportion of this gas in air. Since carbon dioxide is very soluble in water, it follows that if a large surplus of carbon dioxide were introduced into the atmosphere, say, from volcanic

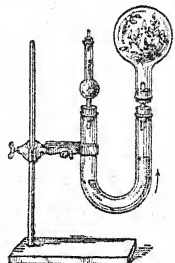


FIG. 290.—Showing the Abstraction of Oxygen from Air during the Respiration of Plants.



FIG. 291.—Assimilation of Carbon Dioxide by Plants.

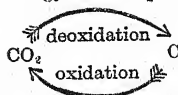
¹ That is, water containing dissolved oxygen, or rather air.

² If the water be too highly charged with carbon dioxide, it may be necessary to remove the carbon dioxide from the oxygen, before testing, by means of a little potassium hydroxide solution. Leaves of the "Common Bird's-foot" were used in the apparatus, Fig. 291

sources, the sea would dissolve a still greater proportion. Hence *the sea must be an important means of regulating the amount of carbon dioxide in the atmosphere.*

It will be noticed that all animal and vegetable life is dependent upon the carbon dioxide \rightarrow carbon reaction, which in turn is dependent upon the sun's energy. There is a kind of closed cycle, an alternation of oxidations and deoxidations:

Sun's energy stored—plants feed



Energy dissipated—animals and plants breathe, etc.

maintained by a continuous supply of energy from the sun. G. Stephenson was wont to call coal "bottled-up sunshine," a saying which, though not strictly accurate, is true in the sense that the sun began the operation thousands of years ago. If the supply of sunlight should cease, the deoxidation of carbon dioxide would stop and the present conditions of life on the earth would come to an end because the available carbon would be transformed into unavailable carbon dioxide.

§ 5. Ignition Point ; Kindling Temperature.

When heated up to the ignition point, a combustible substance absorbs heat in the same way that ice does when its temperature is gradually raised ; but as the ignition temperature is reached, the chemical equilibrium becomes unstable, with a resultant transformation even more striking than that which occurs when ice melts.—S. A. REEVE.

We have seen that the speed of a chemical reaction is usually accelerated by raising the temperature ; and also that in exothermal reactions, heat is evolved in consequence of the chemical action. If finely divided phosphorus be exposed to the air, oxidation commences ; heat is developed by chemical action more quickly than it is conducted away. In consequence, there is a rise of temperature. This accelerates the speed of oxidation and causes a still greater development of heat. This, in turn, still further accelerates the speed of the reaction until, when the temperature has risen to about 60° , the phosphorus bursts into flame. The ignition or kindling temperature is the temperature to which the substance must be heated in order to start combustion or explosion. Phosphorus, at temperatures below its own ignition temperature, does not combine with oxygen fast enough to cause inflammation ; at and above this temperature, the oxidation is attended by combustion. Many substances inflame spontaneously at ordinary temperatures, for instance, phosphorus, boron and silicon hydrides, zinc ethyl, etc. This means that the ignition temperatures of these substances are at or below ordinary atmospheric temperatures. The ignition temperature must not be confused with the temperature of the reaction. The ignition temperature is no more the temperature at which the reaction begins than the boiling point of water is the temperature at which evaporation begins. The explosion of a whole mass of gas, and the ignition of a portion of a gaseous mixture are different phenomena. For pressures of one atmosphere, the ignition temperatures of mixtures of oxygen or air with :

	H_2	CO	CH_4	C_2H_2
Oxygen . . .	580°-590°	637°-658°	536°-700°	520°-630°
Air	580°-590°	644°-658°	650°-750°	520°-630°

Diminishing the pressure half an atmosphere raises the ignition temperature about 5°, and raising the pressure 2 atmospheres lowers the ignition temperature about 30°.—H. B. Dixon and H. F. Coward.

If the heat could be conducted away from the oxidizing phosphorus quickly enough, the temperature of the reacting substance would never reach the ignition point. Substances in vigorous combustion, if cooled below the ignition temperature at the seat of the reaction, may be extinguished. In illustration, a candle flame can be extinguished by placing a helix of copper wire about the flame, whereas, if the helix be first heated, the flame will not be extinguished.

It is not necessary to heat all the reacting system to the temperature of ignition. The heat may be applied locally. A lighted match applied at one point will ignite a barrel of gunpowder; and a small electric spark is sufficient to inflame a vessel of detonating gas. But a sufficient quantity of the mixture must be ignited to develop enough heat to prevent its being conducted away so quickly that the temperature cannot be maintained above the ignition temperature when the stimulus is withdrawn. Thus, sparks below 0.22 mm. in length will not ignite electrolytic gas. The nitrogen and oxygen of atmospheric air can be made to burn in an electric arc with a flame producing nitric and nitrous acids; but the evolution of heat is not sufficiently great to raise the temperature of the gas up to its ignition point. Indeed, the reaction is endothermal. If otherwise, the flame would quickly spread through the atmosphere, and "deluge the world in a sea of nitric and nitrous acids" (W. Crookes). In endothermal reactions, too, the reaction may be started at one point, heat is absorbed from the surroundings, the temperature at the seat of the reaction is reduced, and the speed of the reaction slows down, until finally, the reaction comes to a standstill.

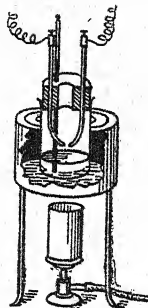


FIG. 292.—Determination of Flash point.

The ignition temperature of oils is an important subject where petroleum oils are used for illumination, cylinder oils in high pressure steam, etc. Grave dangers attend the use of oils with too low an ignition temperature, or, to use the commercial term, too low a flash point. The flash point of an oil is defined as "the temperature at which the oil gives off sufficient vapour to form a momentary flash when a small flame is brought near to surface." The apparatus shown in Fig. 292 is a convenient way of performing the test. A beaker is fitted with a cork having a hole in its centre. A thermometer and two electrodes are also fitted in the cork. The tips of the electrodes are about 5 or 10 mm. apart. The beaker is fixed in a water-bath; the central hole in the cork is covered with a glass plate; the temperature of the bath is gradually raised, and sparks are sent across the wires by connecting the wires with an induction coil and accumulator, when the temperature reaches the flash point, the glass plate is lifted up by a small explosion. The temperature indicated by the thermometer is the flash point of the oil. Slightly different "flash points" are obtained by different methods of applying the test.

It has been found necessary to legislate against the use of kerosene, petroleum, etc., with a flash point below a certain prescribed temperature, which is not the same in different countries: 44°C . or 111°F . is about the average. The flash point of petrol is low, but the ignition temperature is high. It is difficult to determine the exact ignition temperature of gases because the ignition temperature is conditioned by the temperature and pressure of the gas; the rate at which heat is conducted away from the seat of the reaction, etc. Numbers varying from 500° to 845° have been published for the ignition point of electrolytic gas.

§ 6. Explosions.

In ignitable explosives, temperature is in equilibrium with the chemical structure of the molecule so that when more than a certain amount of thermal energy is introduced, per unit of mass, the instability of equilibrium appears most spectacularly in violence of explosion.—S. A. REEVE.

From the preceding section, it will appear that there is no real line of demarcation between rapid combustion and an explosive combustion. The speed of the explosion is determined by the rate at which the process of combustion, initiated at one point, travels through the mixture. Careful measurements indicate that the explosion of gaseous mixtures have this peculiarity: the speed of explosive combustion increases rapidly from its point of origin until it has attained a maximum speed, after that, the explosion travels with the uniform maximum speed. The maximum speed has a specific value for each explosive mixture under definite conditions. This constant is called the velocity of the explosion or detonation wave. If a mixture of gases explodes under such conditions that the explosion wave is inaugurated, the explosion will be far more violent than if it occurred under such conditions that the explosion wave was not established. The explosion wave, at top speed, travels seven or eight times as fast as sound in the same gas. *E.g.*, according to H. B. Dixon (1893):

	$2\text{H}_2 + \text{O}_2$	$2\text{CO} + \text{O}_2$	$\text{CH}_4 + 3\text{O}_2$
Explosion wave	2810	1089	2287 metres per sec.
Sound wave	514	328	345 metres per sec.

According to P. Eitner (1901), if air be mixed with less than 9.45 per cent. or more than 66.40 per cent. of hydrogen an explosion will not take place, but if the amount of hydrogen is between 9.45 and 66.40 per cent. explosion will take place. Similarly, for other gases, H. Bunte and P. Eitner (1904) found:

TABLE LVIII.—PERCENTAGE VOLUME OF GAS REQUIRED FOR AN EXPLOSIVE MIXTURE WITH AIR.

Gas.	Lower explosive limit per cent. of gas.	Upper explosive limit per cent. of gas.	Range between the two limits per cent. of gas.
Carbon monoxide	16.50	74.95	58.45
Hydrogen	9.45	66.40	57.95
Water gas	12.40	66.75	54.35
Acetylene	3.35	52.30	48.95
Coal gas	7.90	19.10	11.20
Ethylene	4.10	14.60	10.50
Marsh gas	6.10	12.80	6.70

The numbers obtained by different methods of ignition, etc., vary a few per cent. The appearance of great accuracy conveyed by the two decimals is illusive. The figures would probably have been as accurate if expressed to the nearest whole number.

Finely divided coal dust or carbon form explosive mixtures with air or oxygen, and serious explosions have occurred in several factories in which finely divided organic matter — charcoal, coal, flour, etc.—suspended in the air has been ignited. The explosion varies in intensity with the amount and fineness of the dust. To illustrate the explosion of mixtures of oxidizable solids with air, spread a half centimetre layer of fine willow charcoal at the bottom of a stout 500 c.c. gas jar, Fig. 293; and direct a stream of oxygen into the jar so that the current of gas stirs up the charcoal and fills the cylinder with a dust-laden atmosphere of oxygen. The mixture explodes when a light is applied at the mouth of the cylinder. Similar explosive results can be obtained with the dust of finely divided metallic iron, aluminium, or zinc. The explosion is here so brilliant that coloured goggles should be used by the operator.

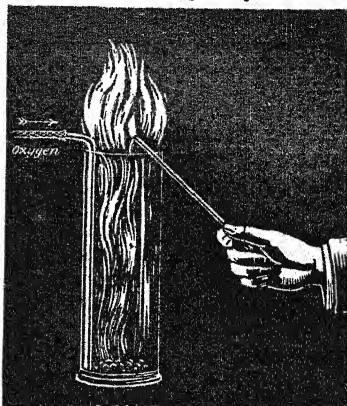


FIG. 293.—Coal-dust Explosion.

§ 7. The Safety Lamp. Flame Caps.

The safety lamp of Davy is a very beautiful application of a great discovery in science to the purposes of humanity.—R. HUNT.

A stout glass tube *B*, Fig. 294, about 3 cm. diameter and 60 cm. long, is fitted at one end *A* with a perforated rubber stopper and bent glass tube; the other end is fitted with a perforated stopper and copper tube *C* about 6 mm. in diameter and 30 cm. long. The tube is clamped to make an angle of about 30° with the horizontal. Fill the tube with coal gas by connecting *C* with the gas supply; light the gas at *C* and remove the stopper at *A*. Air enters at *A* and makes an explosive mixture of air and gas in the tube. The flame at *C* will retreat down the *C* tube and explode the gas in *B*. Repeat the experiment using a copper tube 3 mm. in diameter at *C*. The flame will be extinguished in *C* before it reaches the gas in *B*.¹ That *B* does contain

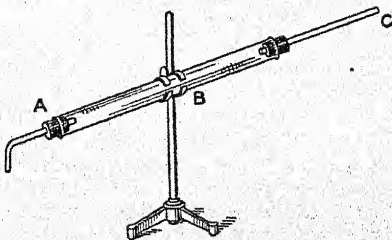


FIG. 294.—Gas Explosion.

¹ A little adjustment is needed in getting the right angle for a successful experiment.

an explosive mixture of air and gas can be shown by applying a lighted taper at *A*. Consequently the flame is cooled below the ignition temperature of the mixture, during its passage down the narrower copper tube; but not with the wider copper tube.

Resuming the subject of ignition temperatures opened on p. 908, fix a piece of wire gauze over an unlighted Bunsen's burner, turn on the gas, bring a lighted taper above the gauze, the gas above the gauze is ignited, but the flame does not pass through the gauze, Fig. 295. Conversely, if the gas be ignited below the gauze, the flame will not pass through, although the passage of inflammable gas through the gauze can be demonstrated by bringing a lighted taper above the gauze. If the gauze be heated red hot, the flame will pass through and burn on both sides.

Experiments not unlike these, led Humphry Davy, in 1815, to the idea of "flame sieves," and finally to the "safety lamp for miners." In Davy's safety lamp, Fig. 296, an oil lamp is surrounded by a cylinder—1½ inches in diameter, and 6 inches high—made of iron wire gauze, having 28 meshes per linear inch. The cylinder is closed at the top with a double layer of gauze *A*, Fig. 296. The lamp is provided with a screw lock to prevent it being opened in the mine, and there is a ring at the top to allow the

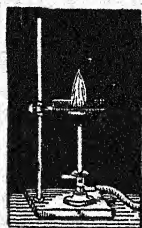


FIG. 295.—Effect of Wire Gauze on Flame.

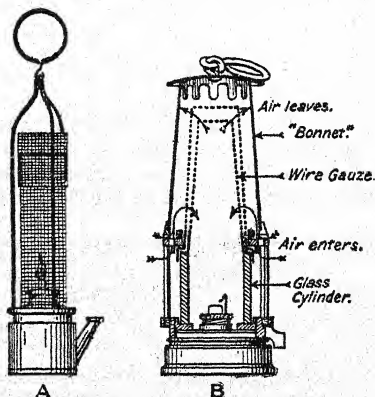


FIG. 296.—Safety Lamps (old and new).

lamp to be carried or suspended. The lamp is ventilated by air passing through the mesh of the gauze near the flame. The air impinges on the flame, and the products of combustion escape through the gauze in the upper part of the cylinder. When such a lamp is taken into an atmosphere containing a mixture of marsh gas and air, the explosive mixture passes through the gauze and is ignited inside the cylinder. The cylinder may (but should not) be heated red-hot by the combustion of the explosive gases inside. The flame, however, is unable to pass through the gauze and ignite the mixture outside the cylinder. The gauze acts as a "flame sieve"; it permits a free passage of gas, air, and light, but it obstructs the passage of the flame. Prior to the invention of Davy's lamp, the risk of explosion in "fiery mines" was very serious; and two notable attempts—by G. Stephenson, and W. R. Clanny—had previously been made to deal with the trouble. The Davy lamp considerably lessened the danger, although it was not fool-proof, and, as might have been anticipated, it sometimes

caused explosions which would otherwise not have occurred because it led miners to take risks in dangerous atmospheres. Mr. Buddle, one of the first users of Davy's lamp, in a Report on Accidents in Mines (1835), stated *inter alia* :

I first tried it in an explosive mixture on the surface, and then took it into a mine ; and, to my astonishment and delight, it is impossible for me to express my feelings at the time when I first suspended the lamp in the mine and saw it red hot ; if it had been a monster destroyed, I could not have felt more exultation than I did.

If the gauze should happen to be heated locally above the ignition temperature of the mixture outside ; or if the lamp is exposed in a strong current of air, say, a ventilating shaft, or a " wave " of air sometimes generated in the operations of blasting, the flame may be driven through the meshes of the gauze. Considerable improvements have been made on the original *Davy's lamp*, A, Fig. 296. The modern forms of safety lamp—e.g. *J. B. Marsaut's lamp*, B, Fig. 296—allow the lamp to be used under conditions of great danger with a minimum risk of igniting an explosive mixture of firedamp and air. The old Davy's safety lamp is obsolete. It gives a very poor illumination owing to the obstruction offered by the wire gauze, and the flame may be forced through the gauze when the current of air exceeds 5 feet per second, and in modern mines, the air in some parts of the rapid ventilating roads may attain 20 to 30 feet per second.

The appearance of the flame of the safety lamp depends upon the illuminant and on the amount of firedamp present in the atmosphere of the mine. The first noticeable effect of firedamp is a slight flickering or " jumping " of the flame ; as the amount of firedamp increases, the flame becomes more and more elongated and smoky until it stretches upwards to the gauze of the lamp as a long thin smoking column. The flame may then leave the wick and burn in the upper part of the gauze with a faint bluish light—called by the miners " corpse-light " ; or the flame may be extinguished.

The peculiar appearance of the flame is due to the formation of an aureole, or flame-cap of burning firedamp which surrounds the flame proper. The flame-cap is transparent and non-luminous by contrast with the bright flame of the lamp itself. To make it visible, the flame, burning

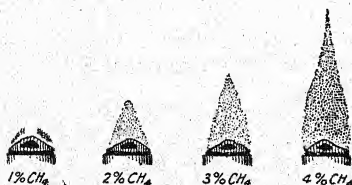


FIG. 297.—Flame-caps—Whalley and Tweedie.

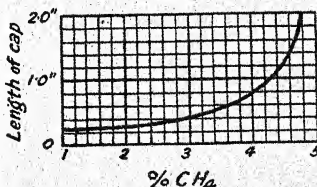


FIG. 298.—Effect of Methane on Length of Flame-caps—Whalley and Tweedie.

from a well-trimmed wick, is " pulled down " as low as possible until a clear blue aureole entirely surrounds a small luminous yellow spot of light, and the flame extends right across the wick. This may be called the standard flame. E. B. Whalley and W. M. Tweedie (1910) found that with " white rose " paraffin oil, the first indications of firedamp occur when about one

per cent. of the gas is present. A very faint cap, without visible tip is then just visible; with two per cent. of firedamp, the tip of the cap is just visible, and the cap has increased in length; the length increases slowly when between 2 and 3 per cent. of firedamp is present; the increase in length is rapid between 3 and 4 per cent. of firedamp; and still more rapid between 4 and 5 per cent. when the flame "spires" up to the top of the gauze. The appearance of the aureole or flame-cap with these different amounts of firedamp is indicated in Fig. 297; and a graph of the experimental results in Fig. 298. It is now possible to form a very fair estimate of the amount of marsh gas in the mine air from the elongation of the flame by using illuminants, etc., standardized in atmospheres containing definite proportions of marsh gas. The method is daily used in coal mines for detecting dangerous amounts of firedamp.

§ 8. The Relation between the Combustible and the Supporter of Combustion.

It has been found convenient in our previous studies to call the burning substance the combustible, and the atmosphere surrounding the burning substance the supporter of combustion. Under ordinary conditions the atmosphere is the enveloping medium, and the terms "combustible" and "incombustible," without further amplification, refer to the burning or non-burning of substances in air. Conventionally too, other gases are

said to be supporters or non-supporters of combustion if they behave towards ordinary combustibles like air. These terms, however, are purely conventional because, from the chemical point of view, it is a matter of indifference whether coal gas burns in air, or air in coal gas. If the atmosphere be coal gas, the flame must be fed with atmospheric air. In this case, the meaning of the conventional terms is reversed, for air is then the combustible, and coal gas the supporter of combustion. There are several ways of showing this reversed or reciprocal combustion, as it is called.

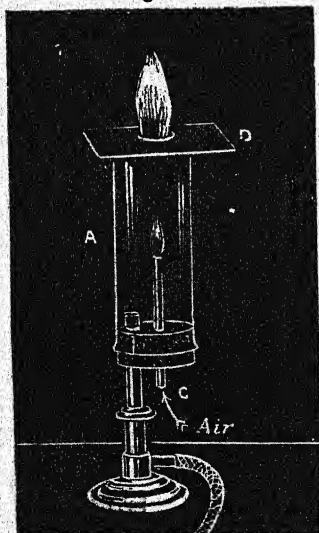


FIG. 299.—Reciprocal Combustion.

An ordinary lamp cylinder—A, Fig. 299—may be closed at one end with a cork perforated to fit over a Bunsen's burner,¹ B, and also with a hole in which a glass tube C—about 10 cm. long and 1 cm. diameter is fitted. The top of the cylinder is covered with a sheet of asbestos D—6 or 7 cm. square, and perforated with a round hole about 2 cm. diameter. The air holes of the Bunsen's burner are closed; the gas is turned on and the opening in the asbestos is closed by laying a piece of cardboard loosely

¹ Or with a glass tube connected with a Bunsen's burner.

on the hole. In a few minutes, the apparatus will be filled with coal gas. Light the gas as it issues from *C*, and simultaneously remove the cardboard which was closing the hole in the asbestos. The flame will pass up the tube drawing the air after it. The upward current of gas causes an upward current of air in the tube *C* which burns with a feebly luminous flame in the atmosphere of coal gas. The excess of coal gas issuing from the opening in the asbestos may be ignited, and the two flames show air burning in coal gas, and coal gas burning in air. By modifying the arrangement, Fig. 299, it is easy to show hydrogen and marsh gas supporters of combustion, and oxygen, chlorine, and nitrous oxide as combustibles.

§ 9. The Structure of Flame.

Burning is not a separation of a special fire-matter, but is a glowing condition of volatile bodies.—J. B. VAN HELMONT (1640).

Is not flame a vapour, fume, or exhalation heated red-hot—that is so hot as to fume? Metals do not flame because they lack a copious fume.—ISAAC NEWTON (1701).

The development of flame during combustion is quite an accidental feature. Iron burning in oxygen gives no perceptible flame. The intense light is due to the incandescent solid. Similar remarks apply generally to the combustion of solids which are not volatilized at the temperature developed during combustion. For this reason, also, copper, carbon,¹ etc., do not usually show any appreciable flame. On the other hand, phosphorus, sulphur, bitumen, fat, wax, etc., burn with a flame because these solids are volatilized at the temperature of combustion.

The flame of most of the combustible gases has quite a characteristic appearance—sulphur burns in air with a lavender blue flame; burning hydrogen is scarcely visible in bright daylight, provided the gas and air are free from dust; carbon monoxide has a rich blue flame, silicon hydride a pale green flame, the cyanogen flame has a delicate pink tinge, acetylene burns with a highly luminous yellowish flame, while marsh gas burns with a feebly luminous yellowish flame.

It will be obvious that when a stream of gas issues from a tube, the gas can only burn at its surface of contact with the air. The shape of the flame, issuing from, say, a circular jet, is due to the fact that as the gas issues from the jet, a ring of gas, so to speak, next to the tube burns first; before any more gas can come in contact with the air and burn, it must rise past the first ring of flame, and each successive layer of gas has to pass higher and higher before it can find the supply of air necessary for combustion. The flame thus assumes the form of a cone. Of course there are no real "rings of gas," but this mode of expression helps us to understand the phenomenon. The upward direction of the flame, even if the gas be directed horizontally or downwards, is due to the increased buoyancy of the hot air in immediate contact with the flame producing a strong upward current of air, and not, as was once suggested, to the "affinity of the heavens" drawing the flame upwards. The particular shape of the flame is also conditioned by the pressure behind the gas issuing from the tube, and on the character of the jet—e.g. the so-called "bat's wings," and "fish-tail" burners.

¹ If carbon burns to carbon monoxide, and then to dioxide, flame may be produced by the burning carbon monoxide.

The interior of the flame which does not come in contact with the air is comparatively cold unburnt gas. Numerous experiments have been devised to show this. For instance :

(1) Cross sections of the flame can be shown by depressing thin sheets of asbestos paper (say, 15 cm. square) for a few moments on the flame of, say, a Bunsen's burner protected from draughts.

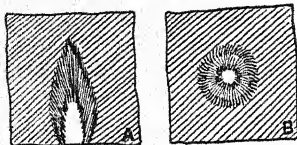


FIG. 300.—Flame Sections.

The hotter portions of the flame where the gases are burning char the paper,¹ producing a dark ring when the paper is held horizontally, Fig. 300, B; and a more or less elongated cone if the paper be held vertically in the flame A, Fig. 300. The particular "flame figure" produced depends on the part of the flame in which the paper is held : N. Teclu (1891).

(2) A narrow glass tube—say 20–30 cm. long—held in the centre of the flame will draw off a portion of the unburnt gas from the interior, and the unburnt gas can be lighted at the end of the tube—Fig. 301.

(3) A pin pushed at right angles through the stem of a match about 3 cm.

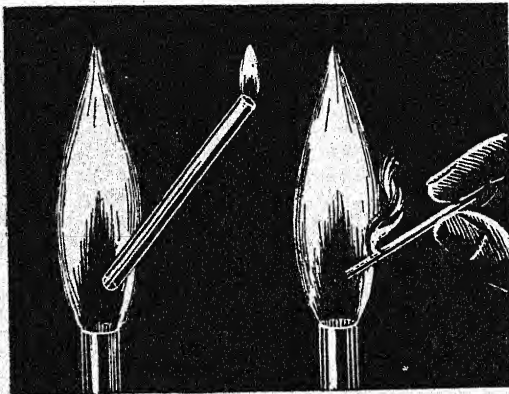


FIG. 301.

Flame Experiments.

FIG. 302.

from the tip, enables the tip of the match to be supported about 3 cm. above the jet of, say, a Bunsen's burner. The gas can be lighted. The match will remain in the centre of the flame without flaming. The experiment can be modified by thrusting the match into the centre of the flame as illustrated in Fig. 302.

(4) The preceding experiment can be modified by connecting a 5–7 cm. funnel with the gas supply as illustrated in Fig. 303. The broad mouth of the funnel is covered with a piece of fine copper or brass wire gauze. A small conical

heap of gunpowder (1½ cm. base) is placed on the middle of the gauze.² The gas is turned on, and a lighted taper slowly depressed from above downwards to the funnel until the gas is ignited. The gunpowder remains on the gauze unconsumed. Ordinary matches can be thrust through the flame and laid on the heap of gunpowder without being ignited.

Single mantled flames.—The nature of the chemical actions which occur in the flame during the burning of gases like hydrogen in air and in chlorine are probably much simpler than is the case with gases like methane, ethylene, and acetylene, for the latter form complex decomposition pro-

¹ Grease or oil is used in making the asbestos paper. Well-glazed white writing paper can also be used. The upper side of the paper can also be dusted with red mercuric iodide. This turns yellow where the flame is hottest. Copper foil also furnishes an interesting modification of the experiment.

² Gunpowder must not be placed on the hot wire gauze, and it must be kept in a compact heap so that no particles are scattered about.

ducts. This is evidenced to some extent by the structure of the hydrogen flame. In the former case, the flame is a cone with a uniform sheath of the burning gas as is illustrated in Fig. 304. The interior cone contains unburnt gas. A section is shown at Fig. 300.

Double mantled flames.—The flame of cyanogen has an inner rose-coloured mantle with an outer bluish sheath. The carbon is oxidized to carbon monoxide in the inner mantle, and to carbon dioxide in the outer mantle. The nitrogen remains practically unoxidized. The flame of hydrogen sulphide in air consists of the usual inner cone surrounded by a lilac coloured mantle which in turn is surrounded by a mantle of deeper blue. The hydrogen sulphide is probably decomposed in the inner part of the flame and immediately burns in the mantles to sulphur dioxide and water. The ammonia flame in oxygen is somewhat similar to the cyanogen flame—the inner mantle is ochre-yellow, and the outer mantle is pale yellow tinged with green. Carbon disulphide resembles hydrogen sulphide, but the carbon burns to carbon dioxide. With carbon disulphide and hydrogen sulphide both constituents are combustible; with ammonia and cyanogen only one constituent is combustible although a trace of the nitrogen may be oxidized.

Burning hydrocarbon gases, burning candles, and burning oils also

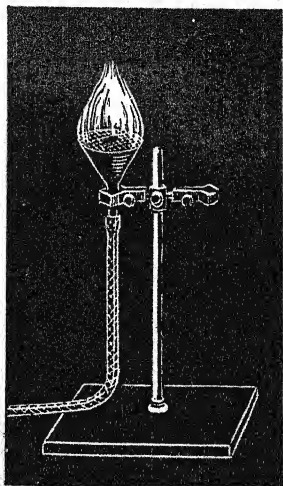


FIG. 303.—Hollowness of Flame.

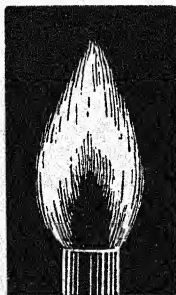


FIG. 304.—Hydrogen Flame.

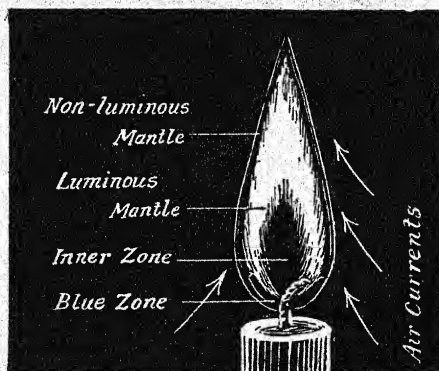


FIG. 305.—Candle Flame.

furnish double sheathed flames—Figs. 305, 307. The inner mantle is bright yellow, more or less opaque and luminous; the outer mantle is but faintly

luminous. In addition, there is a bluish non-luminous zone of incomplete combustion at the base of the flame. With the flame of oil lamps and candles, the hydrocarbons are drawn up the wick by capillary action and vaporized near the top of the wick; in the case of a candle, the wax is melted by the heat of the flame. The wick is so plaited that it curls over and exposes the tip to the air. The tip of the wick is thus gradually burnt, and "snuffing" is not needed.¹

When the flame of gas is gradually lowered, the inner luminous sheath gradually decreases in size, and finally disappears, at the same time, the blue region of incomplete combustion at the base becomes continuous right across the flame, Fig. 307. It is probable that the blue portion represents unburnt gas. Between the flame and the burner there is a

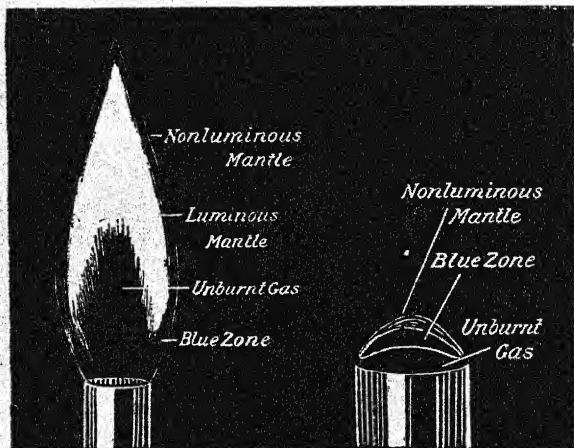


FIG. 306.

FIG. 307.

Flames of Coal Gas.

space containing unburnt gas. The small central cone also contains unburnt gas.

Smithells' flame separator.—A. Smithells has designed an apparatus (1891) which illustrates in an interesting manner the presence of the two mantles in an ordinary coal gas flame. The apparatus consists of two co-axial tubes *A* and *B*, Fig. 308. The outer wider tube is shorter than the inner tube and fitted with a rubber union *C*, and a brass clip to permit the outer tube to be slid up and down. The upper ends of the co-axial tubes are fitted one with a mica *D* and the other with an aluminium *E* cylinder. The narrower tube is clamped over an unlighted Bunsen's burner and a loose packing of cotton wool *F* placed between the burner and the glass tube. Adjust the tops of the co-axial tubes at the same level. Close the air holes of the Bunsen's burner; turn on the gas,

¹ If the wick does not curl, it is liable to project into the inner cone of the flame as the wax is consumed, and thus interfere with the proper combustion of the candle.

light the gas at the top of the outer tube. The gas burns with the usual luminous coal gas flame. Gradually open the air holes of the burner until the flame appears non-luminous. Slide the outer tube upwards, and the two-coned structure of the flame will be obvious, for the outer cone ascends with the wide tube, while the inner cone, after the tube has been pushed upwards about 10 cm., remains burning at the top of the inner tube.

§ 10. The Nature of the Chemical Reactions in Flames.

The chemical reactions which occur in the flames of burning hydrocarbons have not been satisfactorily elucidated, although a little is known. Tapping the gases from different parts of the flame by means of platinum or porcelain tubes connected with an aspirator, is not very satisfactory because of the changed conditions arising when the reactions take place in contact with solids. Hence, the proof that the flame of burning hydrogen sulphide contains free sulphur; and that the luminous mantle of burning hydrocarbons contains free carbon, because a cold dish held in the flame receives a deposit of the respective solids, is quite unsatisfactory. It is conceivable that the hydrogen sulphide flame contains free sulphur, and the hydrocarbon flame free carbon when cold porcelain is held in these flames, but not when this disturbing agent is absent; we therefore prefer demonstrations with methods which do not interfere with the flame itself.

1. The inner cone.—If the flame of coal gas or of a candle be placed between a strong light and a screen, the luminous portion of the flame throws a dark shadow on the screen, and if the flame be made to smoke, the shadow of the luminous portion extends into the shadow cast by the smoke. This test for solid matter in flames is called J. L. Soret's optical test (1875), and is generally taken to prove that free carbon is present in the luminous portion of the flame of a hydrocarbon gas because the flame in the non-luminous portion gives no shadow, nor do the flames of carbon disulphide, phosphorus, and other gases which are known to contain no solid matter. The temperature of the inner cone of unburnt gas is raised by heat radiated from the surrounding mantles of burning gas. It is known that when certain hydrocarbon gases are heated in the absence of air, free carbon and dense hydrocarbons—*e.g.* acetylene—are formed.¹ Acetylene, for instance, is formed. This can be shown by burning air in coal gas

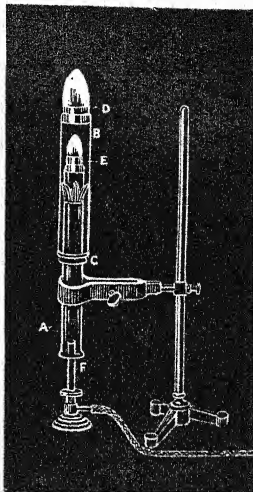


FIG. 308.—Smithells' Experiment.

¹ According to V. B. Lewes (1895) at temperatures between 800° and 1000° part of the ethylene decomposes into acetylene and methane: $3C_2H_4 = 2C_2H_2 + 2CH_4$ along with other more complex hydrocarbons—naphthalene and benzene. Above 1200°, methane, ethylene, acetylene, are decomposed into carbon and hydrogen.

(Fig. 309), so that the ordinary flame is turned inside out, so to speak, when the acetylene, now on the outer surface of the flame, is readily drawn off into the atmosphere of coal gas, and afterwards precipitated by passing the gases through a solution of cuprous chloride (p. 856), as illustrated in Fig. 309.

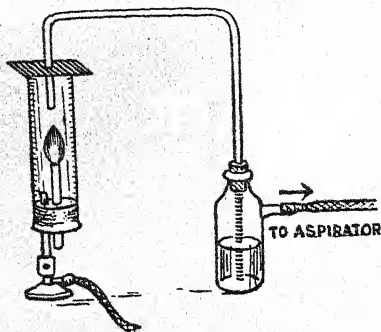


FIG. 309.—Acetylene in Coal Gas Flame.

If the flame be chilled, there is a considerable diminution in the luminosity, so much so that the flame becomes practically non-luminous. This is well illustrated by holding a piece of sheet nickel or iron against a small flame of coal gas burning from a fish-tail burner (Fig. 310). If the opposite side of the sheet of nickel be heated by means of a blowpipe

or large burner, the luminosity of the flame will be restored. On the other hand, if the gas or air be heated, a non-luminous flame becomes distinctly luminous. This can be illustrated by fixing a 15 to 20 cm. tube made of platinum foil over the end of an ordinary Bunsen's burner so as to extend the main tube (Fig. 311). When the platinum tube¹ is heated by the flame of another burner, the non-luminous Bunsen's flame becomes distinctly luminous. Similarly, diluting the coal gas with an inert gas like carbon dioxide leads to a loss in illuminating power. Thus, one of the air holes at the base of a Bunsen's burner may be closed by a cork, and the other fitted with a glass tube connected with a Kipp's apparatus, *K*, generating carbon dioxide, and a drying tube *C* (Fig. 312). By

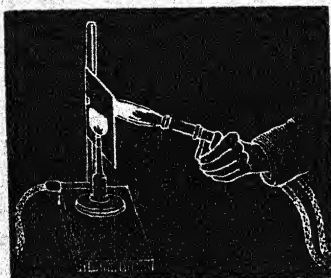


FIG. 310.

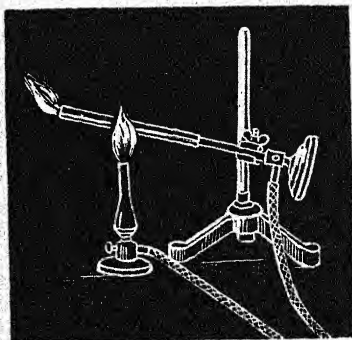


FIG. 311.

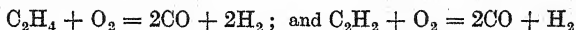
Luminosity of Flames.

gradually increasing the amount of carbon dioxide, the luminosity of the flame can be reduced until it finally becomes blue. If the diluting gas

¹ The platinum tube, of course, should be heated before the test is made to burn off the dust, volatilize the sodium chloride, etc.

be heated, by heating the tube, *B*, the luminosity of the flame is restored. This all seems as if the high temperature flame favours the decomposition of the hydrocarbons in the inner cone into free carbon, etc., and that the free carbon adds to the luminosity of the flame.¹ Hence it is assumed that *in the innermost cone the unburnt hydrocarbons are partly decomposed, at the high temperature of the inner mantle, into free carbon, acetylene, and other hydrocarbons.*

2. The inner mantle.—The result of the analysis of gases in this zone shows that carbon monoxide, hydrogen, carbon dioxide, and water are present.² The little oxygen which can penetrate to the vicinity of the innermost cone oxidizes the hydrocarbons in such a way that carbon monoxide and hydrogen gases still remain unburnt. The idea can be gathered from the equations :



No free oxygen has been detected in this part of the flame. That which

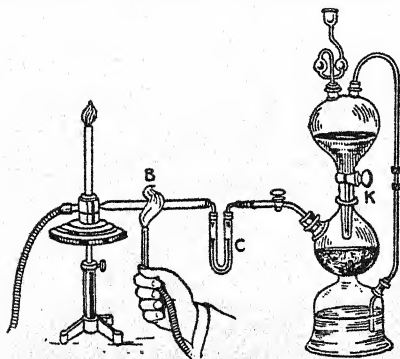


FIG. 312.—Luminosity of Flames.

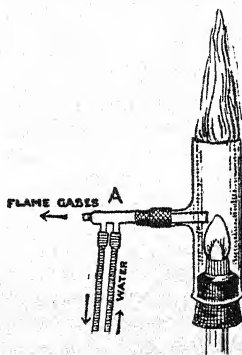
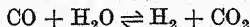


FIG. 313.—Haber's Flame Experiment

does penetrate into this region is used in oxidizing the two unburnt gases. The oxygen distributes itself so that the gases carbon monoxide, water, hydrogen, and carbon dioxide are in equilibrium (see p. 876) :



When this deduction was tested experimentally by F. Haber, the gases were drawn from the space between the inner and outer mantles of the flames of coal gas burning in a Smithells' flame separator fitted with a side tubulure connected with a double-walled water-cooled platinum tube, *A* (Fig. 313). Hence, whatever view be taken of the nature of the chemical process in the inner mantle, the net result of the reactions leads

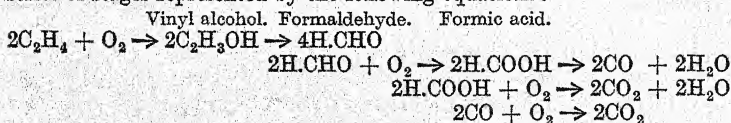
¹ The flame of coal gas burning in oxygen is much reduced in size; the luminosity is increased; and there is not sufficient carbon present to show an appreciable shadow.

² Also nitrogen, of course. This is derived from the atmospheric air. Every one volume of oxygen, it will be remembered, is accompanied by four volumes of nitrogen.

to the conclusion: *In the inner mantle, the hydrocarbons are oxidized to carbon monoxide and hydrogen gases, and these gases are in equilibrium with the carbon dioxide and water, formed at the same time. The relative amount of each gas present is determined by the temperature.*

It was once taught by M. Faraday (1853) and others that when a gas like ethylene— C_2H_4 —is burnt, the hydrogen burns first and the carbon afterwards. As H. B. Dixon has expressed the idea: "There is a race for the oxygen of the air between the two constituents of ethylene, and the hydrogen being the fleetest of the two, gets to the oxygen first and is burnt to water." H. B. Dixon (1893), W. A. Bone and J. C. Cain (1897), and C. M. Kersten (1861) have verified the old observation of J. Dalton (1810): that when a mixture of acetylene or ethylene is detonated with insufficient oxygen to burn all the hydrogen and carbon, carbon monoxide and hydrogen are formed, roughly according to the equation: $C_2H_4 + O_2 = 2CO + 2H_2$; and $C_2H_2 + O_2 = 2CO + H_2$. This certainly makes it appear that the carbon burns to carbon monoxide before the hydrogen. Hence the idea of a selective or preferential combustion. It seems as if nearly all the carbon is first oxidized to carbon monoxide, and the excess of oxygen, if any, divides itself between the carbon monoxide and hydrogen as indicated above.

It is not yet quite clear what really does take place at the temperature of the hydrocarbon flame. There are many reasons for assuming that at rather lower temperatures, the oxidation takes place in a series of "steps" without any liberation of free carbon or free hydrogen at any stage of the process. W. A. Bone heated a mixture of different hydrocarbons with oxygen at a temperature such that neither carbon monoxide nor hydrogen were oxidized with appreciable velocity, nor the steam reduced by carbon. The experiment was arranged so that the gases were circulated alternately through a hot zone and through a washing apparatus for removing condensable or soluble intermediate products of oxidation. The results showed that the oxidation is a complex process. One example will suffice. According to Bone, the oxidation of ethylene probably takes place in the series of stages represented by the following equations:



because all but the vinyl alcohol have been detected among the products of the oxidation. The initial and end stages of the reaction are represented: $C_2H_4 + 3O_2 = 2CO_2 + 2H_2O$. As indicated when discussing slow oxidation, it is highly probable that during the oxidation of hydrocarbons and related materials in slow "natural" decay, p. 904, a very complex series of "humic" compounds are formed before the final products—carbon dioxide and water—appear. Amorphous carbon, when exposed to air in the presence of alkalis at 120° – 150° give red and dark brown colloidal solutions, and is oxidized to alkali formate and oxalate, and finally the carbonate. The experiments of Bone on the oxidation of hydrocarbons at rather higher temperatures, lead us to suppose that the reactions are somewhat simpler; and at the still higher temperature of the flame, the reactions are, in all probability, simpler still.

3. *The outer mantle.*—If everything be properly adjusted there is ample oxygen in contact with the outer mantle to oxidize all the carbon monoxide and hydrogen formed in the inner mantle: $2\text{CO} + \text{O}_2 = 2\text{CO}_2$; and $2\text{H}_2 + \text{O}_2 = 2\text{H}_2\text{O}$. This part of the flame is chilled by the surrounding air. If the gas issues from a jet at an excessive pressure, the air may not have time to oxidize all the carbon (smoke), hydrogen, carbon monoxide, and hydrocarbons. The gases, in consequence, are cooled before oxidation in the outer mantle is completed. Indeed, it is difficult to ensure complete oxidation because the cold air in contact with the flame lowers the temperature of the gases so quickly that a little hydrogen and carbon monoxide sometimes escape oxidation, and hence, traces of these gases are often found among the products of combustion. With a properly adjusted flame, *in the outer mantle, the carbon monoxide and hydrogen formed in the inner mantle are completely oxidized respectively to carbon dioxide and water.*

Surface combustion.—Most surfaces can accelerate the speed of combustion of gases, and often gases will burn in contact with the surface of various substances at temperatures below the ignition point. The contact catalytic action is dependent upon the chemical nature and physical texture of the surface. In the case of coal gas, for instance, the surface appears to be "activated"; the stimulating action of the surface appears to be related with its power of absorbing one or both the reacting gases. The incandescent surface plays a specific rôle in the surface combustion for, as stated by W. A. Bone (1917), "it galvanizes the dormant affinities between the combustible gases and oxygen so that the stately minute of ordinary combustion gives place to the wild intoxication of the Venusberg." In the so-called surface combustion, a mixture of coal gas and air, in the right proportions for perfect combustion, is forced through a porous refractory material like coarse-grained firebrick or bauxite, etc.; the ignited material gets warm and finally incandescent, owing to combustion being confined to the surface of the refractory granules. A large amount of heat is radiated from the surface, and a high temperature can be localized in or concentrated where it is required. The process has been used for cooking and for various industrial operations.

§ 11. The Luminosity of Flames.

A general explanation of the cause of the luminosity of flames is not known. Pure hydrogen burning in dust-free air is almost invisible even in a dark room, although the interior of the flame is slightly green, and the external sheath violet-blue. The hydrogen flame becomes luminous if the gas be passed through a tube containing a plug of cotton wool saturated with benzene. The experiment is usually made by connecting a U-tube with the tube delivering the hydrogen. In Fig. 314 the hydrogen is delivered by a Kipp's apparatus, and the gas is dried in a tower of calcium chloride, *C*. In one leg, *A*, of the U-tube is the plug of cotton wool. Both legs of the U-tube are fitted with jets—preferably fish-tail tips. When all the air is expelled from both legs of the U-tube, light the gas at both jets. The characteristic hydrogen flame appears on one leg, and the luminous "carburetted flame" on the other leg. A non-luminous Bunsen's flame can be made luminous by passing particles of carbon through the flame, say, by rubbing two pieces of charcoal together near the air holes

of the Bunsen's burner; or by causing the air which passes through the air holes to pass over finely divided carbon—say lampblack from turpentine; or the hydrocarbon gases from a burning candle (Fig. 315) can be tapped and led into one of the air holes of the Bunsen's burner adjusted to

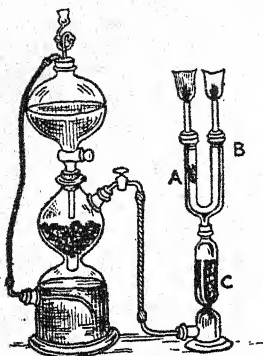


FIG. 314.

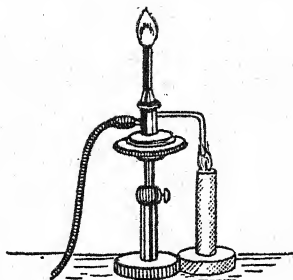


FIG. 315.

burn with a non-luminous flame. When the hydrocarbons are passing through the flame, the flame is luminous. If alcohol is boiled in a flask and the vapour ignited at the mouth of the flask, a pale blue flame is obtained. If a jet of chlorine gas be passed through the flame into the flask, the chlorine gas is ignited, and it burns with a luminous flame (Fig. 316).

The particles of carbon which separate rise into the alcohol flame and make it luminous.

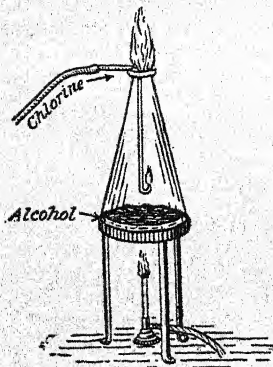


FIG. 316.

Incandescent mantles.—In Welsbach's burner, the flame is non-luminous like an ordinary Bunsen's burner. The Welsbach's mantle is hung so that it is heated by the burning gas. The mantle becomes incandescent. The mantles are made by saturating woven cotton fabrics with a mixture of, say, 99 per cent. thoria and 1 per cent. ceria. The mixed nitrates are dissolved in water. The cotton fabric retains some of the salts in its pores. The cotton is then allowed to dry and burnt off in a stove; at the same time, the nitrates are transformed into oxides. The mantle is hardened and shaped in the blowpipe flame, and the result

is a kind of fragile pottery which retains the form and shape of the original fabric. It would be too fragile for transit. It is therefore dipped in collodion and dried. The collodion is burnt off the first time the mantle is placed in position, and it is no longer to be carried about. The incandescent or light-emitting power of the mantle is somewhat sensitive to variations in the composition of the mixture of earths.

With the above-named mixture, the light has ten times the illuminating power of a flat flame burner using the same amount of gas.

If the proportion of thoria to ceria in the ash of the mantle varies very much from 99 to 98 per cent of thoria and 1 to 2 per cent. of ceria, the luminosity of the mantle is reduced. If a greater proportion of ceria be present, the diminished luminosity is said to be due to the fact that the mantle cannot all be raised by the flame to full incandescence. Thoria and ceria alone give very little light. If a minute quantity of ceria be distributed over the poor heat conducting thoria, the mass of ceria is so small that it can be readily heated to the high temperature needed for producing an intense incandescence. How the *excitant* does its work is not known. The luminosity of the mantle increases with about the fifth power of the temperature.

Davy's solid-particle theory.—As a result of his classical investigations on flame, Humphry Davy (1815) was led to suggest that "the superior light from a stream of coal gas might be owing to a decomposition of a part of the gas towards the interior of the flame where the air is in smallest quantity, and the deposition of solid carbon, which first by its ignition and afterwards by its combustion, increased in a high degree the intensity of the light." And finally, Davy concluded that "whenever a flame is remarkably brilliant or dense (opaque) it may always be concluded that some solid matter is produced in it; on the contrary, when a flame is extremely feeble and transparent, it may be inferred that no solid matter is formed." We have seen how carbon can be produced in a hydrocarbon flame. Davy's generalization is not always satisfactory because burning phosphorus; hydrogen phosphide; carbon disulphide in oxygen or in nitric oxide; hydrogen arsenide; etc., give luminous flames although the temperature of the flame is too high to allow the presence of solid products of combustion. This is confirmed by the fact that these flames give no shadows by Soret's optical test. Consequently, however true the solid-particle theory may be for some hydrocarbon flames, it does not describe all the story.

Frankland's dense hydrocarbon theory.—H. Davy showed that the luminosity of flames is increased by pressure and diminished by rarefaction. The violet blue sheath of the hydrogen flame becomes green, then yellow, orange, and red as the pressure is reduced; and by increasing the pressure, the flame becomes luminous. E. Frankland (1867) found a direct relation between the luminosity of flames and the pressure. Using the flame of hydrogen burning under a pressure of 12 atmospheres, a reduction of the pressure to 3 atmospheres, diminished the luminosity 99 per cent. as illustrated in Fig. 317. Arguing

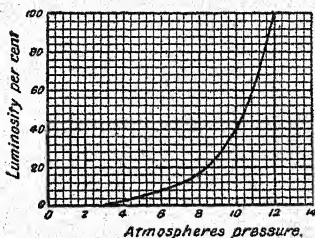


FIG. 317.—Influence of Pressure on Luminosity.

from the fact that luminous flames are known which contain no solids; that dense gases and vapours give flames more luminous than gases of low density; and that feebly luminous flames become luminous when the surrounding atmosphere is compressed, Frankland inferred that the luminosity of ordinary flames such as coal gas is due to the glow of dense hydrocarbons rather than to the

presence of solid particles, Lewes considers that the "dense hydrocarbon" in the flame of coal gas and related gases is acetylene, hence, the so-called *acetylene theory of luminosity*. In the particular case of hydrocarbon flames, as interpreted in the preceding section, it does not appear that this hypothesis is a complete explanation. When the pressure of the atmosphere is altered, the dense hydrocarbons themselves give opaque flames containing solid matter. The phenomenon is complicated by changes in the conditions of equilibrium of the products of combustion, and it is probable that the decomposition of the hydrocarbons in the "innermost parts of the flame" is facilitated by increasing the pressure.

However, it is highly probable that dense vapours, as well as incandescent solids, do produce luminosity. Two distinct effects can be observed in burning hydrogen phosphide: a greenish glow due to the oxidation of the phosphorus which shows best when the phosphide is largely diluted with carbon dioxide; and a yellowish white light best seen when phosphorus burns in air or oxygen. This is no doubt due to the glowing of phosphorus pentoxide which, although in the state of *vapour*, can be made to emit a yellowish white glow when the temperature has attained a certain point. Similar results can be obtained with silicon hydride, but in this case the glow is due to *solid silica*.

Influence of temperature on luminosity.—Just as a variation of pressure modifies the luminosity of burning gases, so does a variation of temperature, as indicated on pp. 920–1. The temperature attained by the combustion of gases in oxygen is greater than in air, owing to the absence of diluting nitrogen. In consequence, combustibles burn far more brilliantly in oxygen than in air. The flame of carbon disulphide in air is nothing like so brilliant as in oxygen; the flame of hydrogen phosphide in air is brilliant, but in oxygen the flame is of dazzling brilliance. Phosphorus burns in chlorine with a far more luminous flame when the chlorine is hot than when cold. Carbon monoxide in oxygen burns with a flame appreciably luminous. Similarly with flames known to contain solid matter. Magnesium and silicon hydride burn far more brightly in oxygen than in air. The effect of increasing the temperature of the gases in a non-luminous Bunsen's burner has been previously discussed.

To summarize: A general explanation of luminosity to cover all cases is wanting. To say that "luminosity is caused by the intense vibratory motions of the molecules induced by the act of chemical union" is not very satisfactory, because it tells us little more than that "flames are luminous because they emit a bright light," until it is shown why the "intense vibratory motions of the molecules" of many non-luminous flames do not give light. In special cases, the luminosity of flames can be traced to the "glowing" produced either by the products of combustion (*e.g.* hydrogen phosphides, silicon hydride) or intermediate products of combustion (*e.g.* hydrocarbon flames). The glowing may be due to the presence of glowing vapours (*e.g.* hydrogen phosphide) or to glowing solid matter (*e.g.* silicon hydride). The luminosity of flames is in general increased by raising the temperature or increasing the pressure.

§ 12. The Bunsen's Burner.

1. *Structure of the burner.*—The gas burner devised by Robert Bunsen, about 1855, will be very familiar to chemistry students. Its construction is

easily understood. Unscrew the burner tube from the base. The burner consists of three parts. (1) The base *A*, Fig. 318, supplied with gas by means of a rubber tube connected with the main. The gas escapes from a small opening in the base, which may or may not have a screw "pin-hole" nipple. By lighting the gas issuing from the base, a long thin pencil of flame is obtained. (2) The burner tube, *B*, has a couple of openings near the base, and these can be closed, partially closed, or opened by turning the air regulator. (3) The air regulator, *C*, is a short cylindrical tube fitted with holes to correspond with the holes in the burner tube. There are numerous modifications; that just described may be taken as typical. Some are provided with an attachment for forcing in air under pressure, forming the so-called *blast burners*. Replace the air regulator and burner tube.

2. How the burner does its work.—If the air-holes are closed, an ordinary luminous gas flame is obtained. If the air-holes are opened, the jet of gas from the small orifice produces a partial vacuum in the neighbourhood of the jet, and, in consequence, air is drawn into the air-holes, and mixes with the gas in the burner tube. The reduction in pressure is conveniently shown by closing one of the air-holes with a piece of gummed paper, and affixing a small manometer, Fig. 319, to the other air-hole by

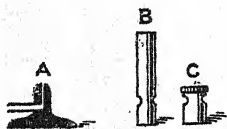


FIG. 318.—Parts of Bunsen's Burner.

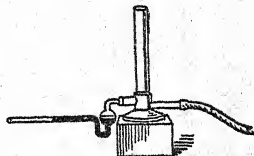


FIG. 319.—Pressure at Air-holes of Bunsen's Burner.

means of a perforated cork.¹ When the manometer is charged with a liquid—coloured to enable it to be seen better—and the gas is turned on, the movement of the liquid towards the burner shows that the air-holes exert a slight suction. When the gas is turned off, the liquid in the manometer returns to its former position.

A certain ratio must exist between the proportion of air and gas in the burner tube in order to get the gas to burn quietly with a blue flame—the so-called Bunsen's flame. This is observed by placing the regulator in several different positions and gradually turning off the gas. When the air-holes are fully open, and the gas is gradually turned off, a point is reached when the flame begins to flicker, and finally "strikes back" afterwards burning at the bottom of the tube. As a matter of fact, the mixture of air and gas burning in the Bunsen's burner is explosive when the gas is burning quietly, the rate at which the flame travels in the explosive mixture of air and gas is less than the rate at which gas is issuing from the burner; when more air or less gas is introduced, the speed at which the explosive flame can travel is increased; when the rate of the explosive flame and the speed of the gas issuing from the burner are nearly equal the flame reaches the unstable condition; any further increase in the amount of air or decrease in the amount of gas gives an explosive

¹ Or solder a piece of brass tubing to the burner.

mixture in which the explosive flame can travel faster than the issuing gas, the flame then "strikes back." This phenomenon is best studied by clamping a long tube—5 or 6 feet long, and about $1\frac{1}{2}$ inches wide—vertically over a Bunsen's burner, Fig. 320. Plug the space between the Bunsen's burner and the glass tube with cotton wool. Close the air-holes of the former, and light the gas (Fig. 320). Gradually remove the cotton wool until a large blue Bunsen's flame is obtained. Then open the air-holes of the Bunsen's burner gradually, and a point will be reached when the flame "strikes back" with a loud though harmless report.

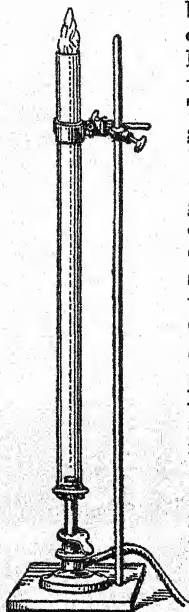


FIG. 320.—Striking back of Bunsen's Flame.

In the Bunsen's burner, the proportion of air to gas is between 2 and 3 volumes of air per one volume of gas, but for complete combustion of the gas, about 6 volumes of air are needed. If such a mixture were sent through an ordinary Bunsen's burner, the flame would "strike back"—for the reasons indicated above.

G. Méker, however, has designed a burner in which such a mixture can be burnt—Méker's burner. From the sectional diagram, Fig. 321, it will be seen that the air-holes are larger than usual, and a deep grid hinders the flame "striking back." Since the gas issuing from the burner has enough air for complete combustion, the flame is practically a "solid cone" of burning gas, and there is no "inner cone of unburnt gas." In consequence,

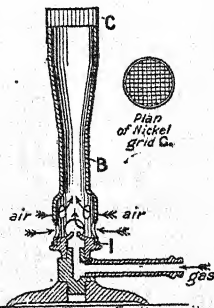


FIG. 321.—Méker's Burner.

the temperature of the interior of the flame is rather greater than the temperature near the outside. The effect of air on the illuminating power of coal gas is illustrated by the following measurements due to Wurtz:

Air	3.0	4.9	11.7	16.2	25.0 per cent.
Loss luminosity	15.7	20.8	41.5	57.5	84.0 per cent.

3. Why is the flame non-luminous?—(a) Oxidation. It was formerly taught that the non-luminosity of the Bunsen's flame was exclusively due to the influence of the admixed oxygen bringing about rapid and complete combustion so that instead of the hydrocarbons decomposing in a series of intermediate stages, they were burnt more directly to carbon dioxide and water. Experiments by V. B. Lewes on the percentage amount of different gases required to produce the non-luminous flame showed that one volume of coal gas required for the non-luminous flame:

Mixture of O : N	1 : 0	1 : 1	1 : 2	1 : 3	1 : 5 (air)
Volume required	0.5	1.0	1.5	2.0	2.3 vols.

This shows that oxygen intimately mixed with the coal gas exerts an oxidizing effect and facilitates rapid combustion. (b) The cooling of the flame by different gases. This, however, is not all the story. Inert gases

like nitrogen, carbon dioxide, and steam produce non-luminous flames. Thus the following percentage volumes of the gases named were needed to make a coal-gas flame non-luminous :

Gas	Air.	Nitrogen.	Carbon monoxide.	Carbon dioxide.
Volumes required	2.3	2.3	5.1	1.3

These facts coupled with the known effect of temperature on the luminosity of flames, show that the air reduces the luminosity of the Bunsen's flame by chilling the reacting gases. (c) The raising of the temperature of the decomposition of coal gas by admixture with inert gases. Lewes has shown, other things being equal, that coal gas, when mixed with nitrogen, is probably more stable when heated in the inner cone than when nitrogen is absent; and further, a different set of decomposition products are obtained when the hydrocarbons are burned at the higher temperature. Further details on the nature of the Bunsen's flame are indicated in preceding sections.

4. Oxidizing and reducing flames.—The outer mantle of the flame, where there is an excess of oxygen, is oxidizing; and the inner region, where combustion is not complete, is reducing. This can be confirmed by holding a piece of copper wire across different parts of the flame. Advantage is taken of this in qualitative analysis where oxidizing and reducing "flame reactions" furnish valuable indications of the composition of a mixture.

5. The temperature of the Bunsen's flame.—The temperature of different parts of the Bunsen's flame, as recorded by a thermocouple inserted in different parts of the flame (C. Féry, 1903) is shown in Fig. 322. The published determinations of the maximum temperature of the Bunsen's flame are somewhat discordant. The discrepancy arises from the presence of several errors in the method used for the determination. The most satisfactory methods show that the maximum temperature is somewhere between 1800° and 1870°. Valuable information respecting the composition of certain mixtures can be obtained by taking advantage of the different volatility of salts, and the different temperatures of the Bunsen's flame. For instance, potassium salts can be volatilized so as to give the flame reaction for potassium, before the sodium has begun to volatilize, if the mixture of the two salts be held in the cooler part of the Bunsen's flame near the burner.

The temperature obtainable by heating a small body in a Bunsen's flame is said to range from 1100° to 1350°; in a Méker's flame, from 1450° to 1500°; in a petrol blow-pipe flame, from 1500° to 1600°; in the oxy-hydrogen flame, about 2000°; in the oxy-acetylene flame, about 2400°; and in the electric arc, about 3500°.

According to Féry (1904), the flame temperature of a Bunsen's flame, fully aerated, is 1871°; and according to V. B. Lewes (1895), the maximum temperature of a luminous flame is 1330°. Hence, it might be concluded

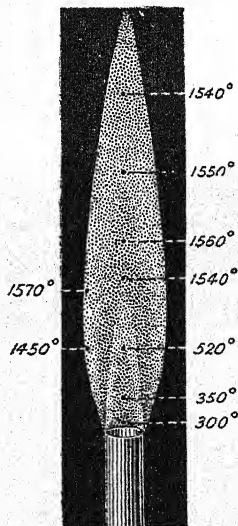


FIG. 322.—Temperatures of Different Parts of the Bunsen's Flame (Féry).

that the Bunsen's flame develops more heat than the luminous flame burning the same amount of gas. This conclusion would be erroneous. Given just sufficient air for complete combustion both flames develop the same amount of heat (Hess' law). The reason the temperature of the luminous flame is less than that of the non-luminous flame is due to the fact that a greater quantity of the heat generated by the combustion of the gas in the luminous burner is lost by radiation. When the luminous flame is placed under a vessel less heat strikes the vessel, because more heat is radiated into space. According to J. Tyndall, the Bunsen's flame radiates 12 per cent. of its heat into space, while the luminous flame radiates 30 per cent. The difference in the heat-radiating power of the two flames can be illustrated by the following experiment. A differential thermo-

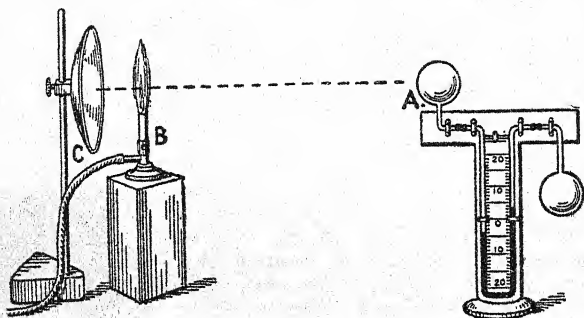


FIG. 323.—Radiation of Heat by Luminous and Non-luminous Flames.

meter, A, Fig. 323, is arranged so that the heat from a Bunsen's flame, B, is concentrated on one of the bulbs by means of a reflector, C. A rise of the temperature of the bulb is shown when the air-holes of the Bunsen's burner are closed and the gas burns with a luminous flame. The difference between the two flames thus depends on the way the heat is distributed. If all the heat in both flames be utilized, the one gives out as much heat as the other. The heat of the Bunsen's flame is more local and concentrated than the heat of the same gas burning with a luminous flame. The Bunsen's flame is therefore best for boiling, etc., where the heat must be concentrated on the bottom of a vessel or as locally as possible. The Bunsen's flame, moreover, is not liable to deposit soot on the bottom of the vessel being heated and so obstruct the conduction of heat from the flame to the vessel. In certain industrial operations the luminous flame often gives better results than the non-luminous flame for heating large enclosures directly by flame. The non-luminous flame heats more where it touches; the luminous flame radiates heat to its surroundings, and this helps to keep the temperature of large enclosures uniform.

Questions.

1. Assuming exterior conditions to be the same, which will probably show the greater explosive effect, a mixture of 500 c.c. of oxygen and 500 c.c. of hydrogen, or a mixture of 300 c.c. of oxygen and 600 c.c. of hydrogen, and why?—*American Coll.*
2. Describe lecture experiments that illustrate—(a) that hydrogen burns

with a luminous flame when benzene is added to it; (b) that a flame "strikes back" through a tube when a sufficient amount of air is mixed with coal gas; (c) that the interior of the inner cone of a round flame is cool; (d) that oxygen will burn in coal gas.—*Cornell Univ., U.S.A.*

3. Give three reasons for the non-luminosity of the flame from a Bunsen's burner. When will such a flame "strike back"? What causes the luminosity of the ordinary gas flame? Why does the amount of carbon dioxide in the atmosphere remain practically the same?—*Cornell Univ., U.S.A.*

4. Describe and explain as fully as you can all the chemical changes which occur in the burning of a deep anthracite coal fire. Explain any differences between such a fire and one of bituminous coal.—*American Coll.*

5. State some of the consequences which would conceivably follow if the percentage of oxygen in the atmosphere were to increase to double the present proportion. Explain your answer, and cite experiments in support of your statements.—*American Coll.*

6. What influence had the phlogistic theory on the progress of chemistry? Why was the anti-phlogistic theory incomplete before the composition of water was known?—*Science and Art Dept.*

7. What is a reducing agent? Name three agents capable of effecting the reduction of ferric chloride in aqueous solution, and explain the chemical changes involved. By what test could the completion of the reduction be ascertained?—*Sheffield Univ.*

8. Write a brief history of the theory of phlogiston with an account of the discoveries which finally led to its rejection.—*Science and Art Dept.*

9. What would be the effect of heating lead chloride to redness, (a) in a current of hydrogen; (b) with potassium carbonate; (c) with charcoal?—*Owens Coll.*

10. Define the terms reducing and oxidizing agents and give three examples of each. Define combustion in (a) the popular and (b) the strictly scientific sense. Will air burn?—*Princeton Univ., U.S.A.*

11. Discuss the theories that have been proposed to account for luminous combustion. Describe the flames of the following substances and their products of combustion, viz.: carbon monoxide, marsh gas, olefiant gas, phosphine, ammonia, hydrogen sulphide, and silicon hydride.—*London Univ.*

12. Describe the sources, preparation, and properties of the earths used in the manufacture of mantles for incandescent lighting. In what proportion are these earths used, and what is the part played by each in producing the intense light emitted by the incandescent surface?—*Board of Educ.*

13. The terms vital air, pure air, dephlogisticated air, and oxygen have at different times been applied to the same substance. Give an account from the historical point of view of the ideas underlying these terms, mentioning the chief chemists whose views they embodied.—*Cape Univ.*

14. Explain the following statement by J. Priestley about 1771, in the light of later investigations: "A sprig of mint vegetable left for a few days in an air vitiated by a burning candle restored the purity of that air sufficiently to allow the candle to burn in it again."

15. J. B. van Helmont "planted a sprig of willow in a vessel of such a soil as appeared to be incapable of yielding it any nutriment; suspended the little willow and its pot in air; fed it on pure water, and noticed that the little plant grew apace, stretching forth its branches and covering itself with leaves." Hence argued van Helmont, that wood, bark, foliage "do lie folded up in some mysterious but not inscrutable manner within the elemental substance water," and all things consist essentially of nothing but water. Where lies the fallacy in van Helmont's conclusion?

16. A small box containing a live guinea-pig is placed on the pan of a delicate balance, and the instrument then exactly counterpoised. If the whole be now allowed to remain at rest, it will soon be seen that there is a distinct diminution of the weight of the box and contents. Explain this fact.—*London Univ.*

17. Describe the construction of a Davy lamp, and indicate the principles upon which its action as a safety-lamp depends.—*London Univ.*

18. Explain how it is that fish can obtain sufficient oxygen for their respiration. What experiment can you suggest for the purpose of proving the correctness of your explanation.—*London Univ.*

CHAPTER XXXIX

THE COMPOUNDS CONTAINING CARBON AND NITROGEN

§ 1. Ferrocyanides and Ferricyanides.

DIESBACH, a colour manufacturer of Berlin, accidentally discovered the colour now known as Prussian or Berlin blue, about 1704, and he mentioned the fact to the alchemist, Dippel, who investigated the subject, and according to Woodward (1724) prepared it by melting dried blood with potash salts and treating the aqueous extract of the mass with ferrous sulphate. J. Brown (1725) showed that animal flesh, and St. Geoffroy (1725) that other animal substances, could be used instead of blood. In 1752 Macquer noticed that when Prussian blue was boiled with an alkali, iron oxide separated from the solution and the mother-liquid contained a substance which separated in yellow crystals. It was then called *phlogisticated potash* and afterwards *prussiate of potash*, it is now known as *potassium ferrocyanide*. K. W. Scheele (1782-5) showed that when distilled with dilute acids, Prussian blue furnishes an acid which he named *prussic acid*. C. L. Berthollet (1787) showed that iron, as well as potash and prussic acid, is an essential constituent of the so-called prussiate of potash.

Potassium ferrocyanide, K_4FeCy_6 .—When nitrogenous refuse (blood, horns, leather scraps, etc.) is charred, and the black mass is ignited with potash and iron filings, something is formed which passes into solution when the mass is lixiviated with water. The aqueous solution on evaporation gives yellow crystals of potassium ferrocyanide with the empirical composition, $K_4FeCy_6 \cdot 3H_2O$. It is convenient to represent the univalent group "CN" by the symbol "Cy," and accordingly the formula is written $K_4FeCy_6 \cdot 3H_2O$. The same salt is obtained from the "spent oxide" of the "purifiers" of gas works which are used (*q.v.*) to remove the sulphur and cyanogen compounds from the gases formed during the distillation of coal. The "spent oxide" is boiled with lime. The soluble calcium ferrocyanide is leached from the mass, and converted into potassium salt by the treatment with potassium carbonate. The resulting potassium ferrocyanide is purified by crystallization. Sometimes the cyanogen compounds are removed from the coal gas before it reaches the purifiers by washing the gas in an alkaline solution with ferrous carbonate in suspension. Sodium ferrocyanide, $Na_4FeCy_6 \cdot 10H_2O$, is produced in an analogous way.

Properties of potassium ferrocyanide.—Potassium ferrocyanide crystallizes in laminated, sulphur-yellow crystals with three molecules of "water

of crystallization," $K_4FeCy_6 \cdot 3H_2O$. The salt is soluble in water, but not in alcohol or ether. The aqueous solution has a bitter taste, but is not particularly poisonous. The salt loses its "water of crystallization" on warming, and the anhydrous salt remains behind as a white powder. The salt is decomposed when heated, forming potassium cyanide, KCy , and an impure iron carbide, generally symbolized FeC_2 , thus, $K_4FeCy_6 \rightarrow 4KC\dot{y} + FeC_2 + N_2$, although it is probably much more complicated than this. When warmed with *dilute* sulphuric acid, hydrogen cyanide (*q.v.*) is formed; with *concentrated* sulphuric acid, carbon monoxide is evolved: $K_4FeCy_6 + 6H_2O + 6H_2SO_4 = 2K_2SO_4 + FeSO_4 + 3(NH_4)_2SO_4 + 6CO$. Potassium ferrocyanide decomposes when ignited out of contact with the air and forms a complex mixture of potassium cyanide, iron carbide, etc. J. A. Müller discovered potassium carbonyl ferrocyanide, K_3FeCy_5CO —in which one carbonyl radicle, CO , takes the place of one Cy radicle in potassium ferri-cyanide—in the mother-liquid obtained when potassium cyanide is prepared by lixiviating roasted potassium ferrocyanide. It is also made by heating carbon monoxide with potassium ferro- or ferri-cyanide in a sealed tube. The reaction is represented: $K_4FeCy_6 + CO + 2H_2O = NH_3 + HCOOK + K_3FeCy_5CO$. The free acid has been isolated in colourless scaly crystals.

Ferrocyanic acid, H_4FeCy_6 .—When a saturated solution of potassium ferrocyanide, freed from dissolved air by boiling, is treated with concentrated hydrochloric acid in the cold, a white crystalline powder called ferrocyanic acid is formed, it has the empirical formula, H_4FeCy_6 . Ferrocyanic acid turns blue on exposure to the air owing to the partial decomposition of the salt and the formation of Prussian blue (*q.v.*), and when heated in *vacuo* at about 300° it furnishes a pale yellow powder with the empirical formula, $FeCy_2$; this powder is probably not a simple cyanide, but rather a ferrous ferrocyanide (*vide infra*). When silver ferrocyanide is boiled with ethyl iodide in alcoholic solution, ethyl ferrocyanide is formed as a crystalline solid. The lowering of the freezing point of aqueous solutions agrees with the formula $(C_2H_5)_4FeCy_6$.

Sodium nitroprusside, $Na_2FeNOCy_5 \cdot 2H_2O$.—When, say, four grams of powdered potassium ferrocyanide are boiled for half an hour with 4 c.c. of concentrated nitric acid diluted with its own volume of water, and the cold solution made alkaline with sodium hydroxide, ruby-red crystals of sodium nitroprusside are obtained. Here the radicle NO takes the place of two Na atoms, and one Cy radicle in sodium ferrocyanide: $3NaCy \cdot FeCy_3 + NO = NaCy + Na_2FeCy_5NO$. When sodium nitroprusside is treated with alkalis, it furnishes sodium ferrocyanide, ferrous hydroxide, etc.: $6Na_2FeCy_5NO + 14NaOH = Fe(OH)_2 + 5Na_4FeCy_6 + 6NaNO_2 + 6H_2O$, showing that the iron is probably present in the ferrous condition. The salt is sometimes called *sodium nitroferrocyanide*. A solution of sodium nitroprusside is sometimes used as a test for sulphides since it gives a deep violet coloration with soluble sulphides. When reduced in alkaline solution (sodium amalgam in water), ammonia and sodium ferrocyanide are formed; when reduced in neutral solution (zinc-copper couple), hydrogen cyanide and sodium ferrocyanide are formed; and when reduced in acid solution (sodium amalgam in dilute sulphuric acid), hydrogen cyanide, ammonium sulphate, and $FeNaFeCy_6$ are formed. Nitroprussic acid, H_2FeCy_5NO , has been obtained in dark-red monoclinic

crystals by treating the silver salt with hydrochloric acid, or the barium salt with dilute sulphuric acid.

Potassium ferricyanide, K_3FeCy_6 .—If an aqueous solution of potassium ferrocyanide be treated with oxidizing agents like chlorine, bromine, nitric acid, hydrogen peroxide, etc., it acquires a dark reddish colour, and crystals of potassium ferricyanide separate when the solution is concentrated by evaporation: $2K_4FeCy_6 + Cl_2 = 2KCl + 2K_3FeCy_6$. The potassium ferricyanide is separated from potassium chloride by re-crystallization. If an excess of chlorine be passed into a solution of potassium ferrocyanide, the solution deposits a greenish precipitate—called *Prussian green*. The oxidation of the ferrocyanide to ferricyanide is now conducted by the electrolysis of slightly alkaline solutions—hydrogen and potassium hydroxide are formed at the cathode: $2K_4FeCy_6 + 2H_2O \rightarrow 2K_3FeCy_6 + 2KOH + H_2$. The ferricyanide is separated by crystallization. Potassium ferricyanide, also called *red prussiate of potash*, is a mild oxidizing agent in alkaline solutions, for it oxidizes “reducing agents” like sodium thiosulphate, hydrogen sulphide, etc., re-forming potassium ferrocyanide: $4K_3FeCy_6 + 4KOH = 2H_2O + 4K_4FeCy_6 + O_2$. Thus, an aqueous solution of the ferricyanide is converted into ferrocyanide by potassium or sodium amalgam. Potassium ferricyanide, also called red prussiate of potash, is rather more soluble in water than the ferrocyanide: thus, 100 grams of water at 0° dissolve 31 grams of potassium ferricyanide, and 13 grams of ferrocyanide; at 10° these numbers are respectively 36 and 20. Potassium ferricyanide is not soluble in alcohol. When a saturated solution of potassium ferricyanide is treated with concentrated hydrochloric acid, in the cold, reddish-brown, acicular crystals of tribasic ferricyanic acid, H_3FeCy_6 , separate from the solution concentrated by evaporation in vacuo. It has been claimed by Locke and Edwards that there is a green modification of potassium ferricyanide which they call the β -form; the ordinary form being the α -salt. O. Hauser and E. Biesalsky consider that the alleged β -form of potassium ferricyanide is a colloidal solution of Prussian blue in a solution of the ordinary red crystals.

Various salts of ferrocyanic and ferricyanic acids have characteristic colours, and consequently, potassium ferrocyanide and ferricyanide—particularly the former—are used in qualitative analysis.

TABLE LIX.—PROPERTIES OF FERRO- AND FERRI-CYANIDES.

	Ferrocyanides added to	Ferricyanides added to
Ferric chloride .	Deep blue, precipitate of Prussian blue, insoluble in hydrochloric acid, soluble in oxalic acid.	No precipitate in neutral solutions, but the solution is coloured green or blue.
Ferrous chloride .	Bluish-white precipitate which rapidly darkens on exposure to air, or by adding a drop of bromine.	Deep blue precipitate of Turnbull's blue.
Copper sulphate .	Reddish-brown precipitate.	Yellowish-green precipitate.
Zinc sulphate .	White precipitate.	Orange precipitate.
Silver nitrate .	White precipitate.	Reddish-brown precipitate.

Ferric ferrocyanide, $\text{Fe}_4(\text{FeCy}_6)_3$.—This compound is also called **Prussian blue**, or *Berlin blue*. It is formed, as indicated in the preceding table, when a solution of potassium ferrocyanide is added to a solution of a ferric salt. It is insoluble in hydrochloric acid, but soluble in oxalic acid, forming a deep blue solution. When heated with concentrated sulphuric acid Prussian blue yields hydrocyanic acid; and when boiled with alkaline hydroxides, ferric hydroxide is precipitated, and alkaline ferrocyanide remains in solution. Besides the "insoluble" Prussian blue, a *soluble or colloidal Prussian blue* is formed when a ferric salt is added to a solution of potassium ferrocyanide, or a ferrous salt to a solution of potassium ferricyanide. By the addition of salt to the solution, the "soluble" Prussian blue is coagulated or "salted out," and the precipitate is then "insoluble" Prussian blue.

Ferrous ferricyanide, $\text{Fe}_3(\text{FeCy}_6)_2$.—When potassium ferricyanide is added to neutral or acid solutions of ferrous chloride, a dark blue precipitate of ferrous ferricyanide, also called "**Turnbull's blue**," is formed. When potassium ferrocyanide is added to a ferrous salt, **ferrous potassium ferrocyanide, $\text{FeK}_2\text{FeCy}_6$** , or **ferrous ferrocyanide, Fe_2FeCy_6** , is formed.

§ 2. Hydrocyanic Acid and the Cyanides.

Molecular weight, HCy or $\text{HCN} = 27$; melting point, -15° ; boiling point, 26.1° ; specific gravity of liquid, 0.6969 at 18° ; and vapour density 0.944 (air unity) at 19° .

Potassium cyanide, KCy .—Potassium cyanide was formerly made by heating potassium ferrocyanide either alone or mixed with potassium carbonate in an iron crucible to a red heat: $\text{K}_4\text{FeCy}_6 + \text{K}_2\text{CO}_3 = 5\text{KCy} + \text{KCyO} + \text{Fe} + \text{CO}_2$. The mass was lixiviated with water, and the solution evaporated to dryness, fused, cast into sticks, and sold as potassium cyanide, although it always contains some potassium cyanate, KCyO . Potassium cyanide is extensively used in electroplating, gilding, the extraction of gold from quartz, and in photography. Fused potassium cyanide is a powerful reducing agent, and it liberates metals from their oxides, and is at the same time converted into potassium cyanate (*q.v.*): $\text{SnO}_2 + 2\text{KCy} = 2\text{KCyO} + \text{Sn}$. Hence potassium cyanide is used in metal-lurgy and in analytical work. When heated alone, potassium cyanide fuses without decomposition in the absence of air, but if air be present, it is partially converted into potassium cyanate (*q.v.*). Potassium cyanide is soluble in water and in hot alcohol. The aqueous solution is very unstable, p. 387, and when boiled with water, it slowly decomposes, forming ammonia and potassium formate, H.CO.OK .

Manufacture of cyanides.—Cyanides can be manufactured cheaply in several ways. By fusing potassium ferrocyanide with sodium, all the cyanogen, Cy , is converted into cyanide: $\text{K}_4\text{FeCy}_6 + 2\text{Na} = 2\text{NaCy} + 4\text{KCy} + \text{Fe}$. The iron can be separated from the fused mass, and a mixture of sodium and potassium cyanides remains. The mixture can be used for the extraction of gold, etc. If ammonia be passed over heated sodium, sodamide is formed: $2\text{NH}_3 + 2\text{Na} = 2\text{NH}_2\text{Na} + \text{H}_2$, as indicated on p. 654; and if the fused mass be run over red-hot carbon, sodium cyanide is formed: $2\text{NaNH}_2 + 2\text{C} = 2\text{H}_2 + 2\text{NaCy}$.

A. Frank and N. Caro patented a process in 1895 for the fixation of

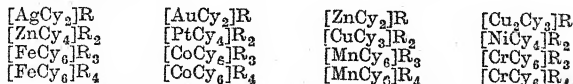
the nitrogen of atmospheric air by heating calcium or barium carbides in an atmosphere of nitrogen between 1000° and 1100° . Nitrogen, it will be remembered, is a by-product in Linde's process for separating oxygen from liquid air. The absorption of nitrogen by the carbides commences about 700° , but the reaction is incomplete; at 1100° , however, the absorption is practically complete, and calcium cyanamide, CaCN_2 , that is, $\text{Ca} = \text{N} - \text{Cy}$, is formed: $\text{CaC}_2 + \text{N}_2 \rightarrow \text{CaCN}_2 + \text{C}$. A large amount of heat is evolved at the same time. The mixture of carbon and calcium cyanamide so obtained is known in commerce as "nitrolime" or "Kalkstickstoff." When in contact with water, calcium cyanamide forms dicyandiamide $(\text{CyNH}_2)_2$, thus: $2\text{CaCN}_2 + 4\text{H}_2\text{O} = 2\text{Ca}(\text{OH})_2 + (\text{CyNH}_2)_2$; and when treated with superheated steam, calcium cyanamide forms calcium carbonate and ammonia: $\text{CaCN}_2 + 3\text{H}_2\text{O} = \text{CaCO}_3 + 2\text{NH}_3$. Calcium cyanamide is used as a fertilizer, and in the manufacture of cyanides, for if calcium cyanamide be melted with a suitable flux—sodium chloride or carbonate—sodium cyanide is produced: $\text{CaCN}_2 + \text{C} = \text{CaCy}_2$; and $\text{CaCy}_2 + \text{Na}_2\text{CO}_3 = \text{CaCO}_3 + 2\text{NaCy}$.

Complex cyanides.—When simple and complex cyanides are boiled with water holding yellow mercuric oxide in suspension, mercuric cyanide, HgCy_2 , is formed; thus, with potassium ferrocyanide: $\text{K}_4\text{FeCy}_6 + 3\text{HgO} + 3\text{H}_2\text{O} = \text{Fe}(\text{OH})_2 + 4\text{KOH} + 3\text{HgCy}_2$. The decomposition of the cyanides by mercuric oxide is utilized in quantitative analysis for the separation of cyanides. Silver cyanide is formed as a white insoluble powder when potassium cyanide is added to a soluble silver salt. The precipitate is soluble in excess, forming a complex potassium argentocyanide, KAgCy_2 . If nitric acid be added to the solution, silver cyanide is precipitated. This complex cyanide is used in electroplating. This reaction is used for the *volumetric determination of potassium cyanide* by J. von Liebig's process. A standard solution of silver nitrate is added to the cyanide solution until a precipitate is just formed. The burette is then read; and the amount of potassium cyanide corresponding with the silver solution dropped from the burette is computed from the equation: $2\text{KCy} + \text{AgNO}_3 = \text{KAgCy}_2 + \text{KNO}_3$. Each atom of silver corresponds with two molecules of potassium cyanide. Any further addition of the silver nitrate will decompose some of the potassium argenticyanide and form a precipitate. The *volumetric determination of silver* can be effected by reversing the process, namely, by titrating a solution of a silver salt of unknown strength with a standard solution of potassium cyanide.

Cupric cyanide, CuCy_2 , as in the analogous case of cupric iodide, is probably formed when potassium cyanide is added to an aqueous solution of, say, cupric acetate, but the salt is so unstable that it immediately passes into cuprous cyanide, CuCy . Cuprous cyanide dissolves in potassium cyanide, forming a similar complex salt: $\text{CuCy} + 3\text{KCy} \rightleftharpoons \text{K}_3\text{CuCy}_4$. The potassium cuprocyanide so obtained is sufficiently stable to be unaffected by hydrogen sulphide in neutral or alkaline solution; the cadmium complex cyanide, K_2CdCy_4 , is decomposed under the same conditions. A common method of separating copper from cadmium depends upon this fact.

The cyanides are remarkable in forming a series of complex cyano-salts remarkable for their stability. Potassium ferro- and ferri-cyanides, and potassium silver cyanide have been previously studied. If the method

worked out for the graphic formula of ammonium chloride were applied consistently it might be inferred the iron in the ferrocyanides is 10-valent, and 9-valent in the ferricyanides. Very little is known about the relative position of the atoms in these molecules. The complex cyanides are conveniently formulated according to Werner's scheme (R univalent), p. 792, in illustration :



In sodium nitroprusside, or sodium nitroferrocyanide, we have $[\text{Fe}(\text{NO})\text{Cy}_5]\text{Na}_2$. The union of ferrous and potassium cyanide in the case of potassium ferrocyanide, and of ferric and potassium cyanide in the case of potassium ferricyanide, must involve a profound change in the molecules concerned. The iron ceases to act as a basic element, but becomes an integral part of the acidic radicle. Potassium ferricyanide is not therefore a double salt, $\text{FeCy}_3 \cdot 3\text{KCy}$; and potassium ferrocyanide is not a double salt, $\text{FeCy}_2 \cdot 4\text{KCy}$, since the iron cannot be separated by precipitation as is the case with the iron in ordinary ferrous and ferric salts; and aqueous solutions of potassium ferrocyanide, according to the ionic hypothesis, contain the quadrivalent ion, FeCy_3''' ; and the ferricyanides, the trivalent ion FeCy_6''' , because the iron appears at the anode, not the cathode, during electrolysis.

Hydrogen cyanide, HCN .—Hydrogen cyanide is made by distilling a mixture of powdered potassium cyanide with a mixture of equal volumes of sulphuric acid and water; if concentrated acid be used, a considerable amount of carbon monoxide is evolved. The vapour is passed through a U-tube containing calcium chloride to remove the water. The dry hydrogen cyanide is led through a U-tube surrounded by ice; and the gas condenses to a colourless liquid, Fig. 324. The gas is also made by passing dry hydrogen sulphide over dry mercury cyanide and condensing the vapour to a liquid as before. Pure hydrogen cyanide is one of the most deadly poisons known, and hence great care must be taken in experiments with hydrogen cyanide, and, indeed, with cyanides generally. The liquid boils at 26.5° , and freezes at -15° to a white solid. It dissolves in water in all proportions, and the solution—called hydrocyanic acid—

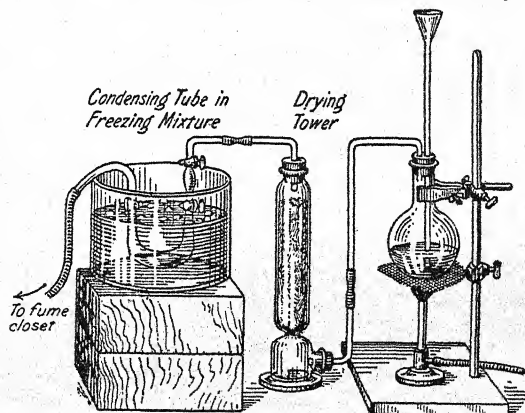


FIG. 324.—The Preparation of Anhydrous Hydrogen Cyanide.

has the smell of bitter almonds. K. W. Scheele discovered hydrocyanic acid in 1782; and made it by heating sulphuric acid with Prussian blue, hence the old name, *prussic acid*; and I. von Ittner made the anhydrous compound, HCN, in 1809. J. L. Gay-Lussac established its composition in 1811. A 10 per cent. aqueous solution is often called "prussic acid," and a 2 per cent. solution is used in medicine. The ordinary aqueous acid can be made by the distillation of, say, 0.5 gram of potassium ferrocyanide with 100 c.c. of 10 per cent. sulphuric acid until 10 c.c. of a dilute aqueous solution of the acid has collected in the receiver. The latter process is also used on a large scale. Hydrocyanic acid is used in the fumigation of ships to destroy rats and vermin, and of railway sleeping cars for bugs and lice. It is also used in spraying fruit trees to destroy certain pests.

Hydrocyanic acid is monobasic, and the salts, as indicated above, are called cyanides. Hydrocyanic acid is one of the weakest of acids, and this corresponds with its low electrical conductivity. Some ammonium cyanide is formed when ammonia is passed over red-hot charcoal; and when a series of electric sparks are passed through a mixture of acetylene and nitrogen. When chlorine gas is passed into hydrocyanic acid, a colourless liquid called "liquid" cyanogen chloride, with the empirical formula CyCl , is produced: $\text{HCy} + \text{Cl}_2 = \text{HCl} + \text{CyCl}$. This polymerizes on standing and forms "solid" cyanogen chloride, or cyanuric chloride, Cy_3Cl_3 . The action of bromine on a metallic cyanide or on hydrocyanic acid furnishes cyanogen bromide, CyBr . This substance sublimes between 60° and 65° , forming transparent crystals. When treated with a well-cooled aqueous solution of sodium trinitride, Na_3N , p. 672, freshly prepared cyanogen bromide gradually dissolves. If the solution be extracted with ether, and the ethereal solution be evaporated in a current of dry air, a colourless liquid is obtained which soon crystallizes. The crystals have the empirical composition CN_4 , and they appear to be cyanogen trinitride or cyanogen hydrazoate, $\text{N}_3\text{—C}\equiv\text{N}$, formed by the reaction: $\text{N}_3\text{Na} + \text{CyBr} = \text{NaBr} + \text{N}_3\text{Cy}$. The compound is explosive; decomposes at 70° ; melts between 35.5° and 36° ; dissolves in water, and the aqueous solution gradually hydrolyzes: $\text{CyN}_3 + 2\text{H}_2\text{O} = \text{N}_3\text{H} + \text{CO}_2 + \text{NH}_3$. The compound gradually polymerizes on keeping.

When hydrocyanic acid is heated with mineral acids, or when potassium cyanide is boiled with water, formic acid, or rather ammonium formate, is produced: $\text{HCy} + 2\text{H}_2\text{O} = \text{H.CO.OH} + \text{NH}_3$. When ammonium formate is distilled with some dehydrating agent—say, phosphorus pentoxide—the formic acid is resolved into hydrogen cyanide: $\text{H.COONH}_4 = \text{HCy} + 2\text{H}_2\text{O}$. These facts, together with much evidence discussed in text-books of organic chemistry, show that the hydrogen atom in hydrogen cyanide is probably united directly with the carbon atom and not with the nitrogen atom, and that the formula of hydrogen cyanide is $\text{H—C}\equiv\text{N}$. The fact that the hydrogen of hydrocyanic acid can be displaced by the metals corresponds with the close analogy between hydrocyanic and hydrochloric acids. There are some reasons for supposing that there are two series of compounds derived from an acid with the empirical formula, HCN; the one set called the cyanides or nitriles corresponds with $\text{H—C}\equiv\text{N}$; and the other, called isocyanides or isonitriles, with $\text{H—N}\equiv\text{C}$. These compounds are discussed in organic chemistry.

§ 3. Cyanogen.

Molecular weight, Cy or C_2N_2 , 52.02; Cy behaves as a univalent radicle; melting point, -34.4° ; boiling point, -20.7° ; critical temperature, 128.3° ; critical pressure, 59.6 atmospheres; vapour density (air unity), 1.805.

Plutarch says that the leaf of the peach tree is the symbol of the god of silence, and that the Egyptians called this deity *Moth*, a term which, the scholars tell us, becomes in Hebrew the word for death. It appears that during the terrible ordeal of initiation into the sacred art of the Egyptians, the candidate was forbidden to speak; an ancient papyrus preserved in the Louvre Museum (Paris) says: "It was forbidden to speak the name of JAO for fear of death by the peach." The punishment was death, and death by poisoning with water distilled from the leaves of the peach tree. The distillate from the leaves of the laurel, bitter-almond, peach, etc., contains hydrocyanic acid. The leaves do not contain the

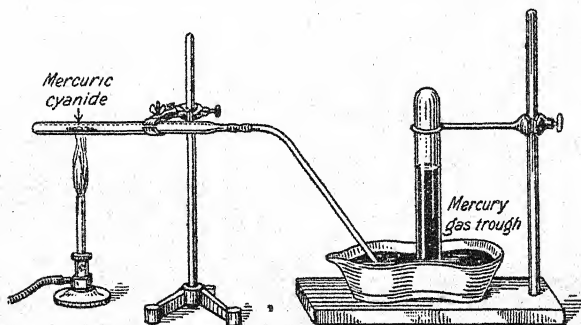


FIG. 325.—The Preparation of Cyanogen.

acid ready formed, but rather a glucoside—called *amygdalin*—which splits up into sugar, oil of bitter almonds, and hydrocyanic acid. The so-called "laurel-water" and "bitter almond-water" of pharmaceutical chemistry contain a little prussic acid. The "jealousy water" employed by the old Egyptians and Hebrews for swiftly poisoning people guilty of certain crimes, was a similar preparation.

Morren (1859) and Berthelot (1879) showed that cyanogen can be obtained directly from its elements by sparking carbon electrodes in nitrogen gas—Fig. 269; and cyanides are formed as indicated above, by heating the two elements in contact with a metal—preferably of the alkalis or alkaline earths. The endothermal reaction is symbolized: $2C + N_2 = Cy_2 - 82 \text{ Cals.}$

Cyanogen, Cy_2 , is a gas made by heating mercuric or silver cyanide in a hard glass tube: $HgCy_2 = Hg + Cy_2$. The gas is best collected over mercury, Fig. 325. The yield of cyanogen is much less than the theoretical owing to the simultaneous formation of a peculiar dark brown powder called paracyanogen. This substance appears to be a polymer of cyanogen, Cy_n , because if continuously heated at about 850° , it furnishes ordinary cyanogen: $Cyn \rightleftharpoons \frac{1}{n}nCy_2$. If the mercuric cyanide be mixed with a little

mercuric chloride, the cyanogen comes off at a lower temperature, and less paracyanogen is formed: $\text{HgCy}_2 + \text{HgCl}_2 = 2\text{HgCl} + \text{Cy}_2$. The gas is also obtained by heating a mixture of mercuric chloride and potassium cyanide; or an aqueous solution of copper sulphate and potassium cyanide. When a mixed solution of potassium cyanide and copper sulphate is warmed, a yellow precipitate of cupric cyanide, CuCy_2 , is first formed, and this is immediately decomposed into cuprous cyanide, CuCy , and cyanogen: $4\text{KCy} + 2\text{CuSO}_4 = 2\text{K}_2\text{SO}_4 + 2\text{CuCy} + \text{Cy}_2$.

Cyanogen is a colourless poisonous gas with a faint odour which reminds some people of the smell of peaches. Cyanogen burns with a violet-coloured flame, forming carbon dioxide and nitrogen. It condenses to a liquid under a pressure of four atmospheres at ordinary temperatures, and at ordinary pressures it furnishes a liquid boiling at -20.7° . The liquid freezes to a white solid melting at -34° . The gas has a vapour density of 52 ($\text{H}_2 = 2$), which corresponds with the molecule C_2N_2 . The graphic formula is probably $\text{N}\equiv\text{C}-\text{C}\equiv\text{N}$. Cyanogen dissolves readily in water. The aqueous solution deposits a peculiar brown flocculent mass—azulmic acid—on standing; ammonium oxalate, hydrogen cyanide, and carbon dioxide are formed at the same time. Cyanogen unites directly with the alkali metals, forming cyanides. The name cyanogen is derived from *κύανος* (cyanos), dark blue; and *γεννάω* (gennao), I produce. Cyanogen was isolated by J. L. Gay-Lussac in 1815. Small quantities of cyanogen occur in blast furnace gases.

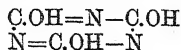
Cyanogen bears some analogy with chlorine. For instance, when it is passed into a solution of potassium hydroxide, it forms potassium cyanate, KOCy , and potassium cyanide, KCy , thus: $\text{Cy}_2 + 2\text{KOH} = \text{KCy} + \text{KCyO} + \text{H}_2\text{O}$. Chlorine under similar conditions forms potassium hypochlorite, KOCl , and potassium chloride, KCl , thus: $\text{Cl}_2 + 2\text{KOH} = \text{KCl} + \text{KOCl} + \text{H}_2\text{O}$. With hydrocyanic acid, the reaction is $\text{HCy} + \text{KOH} = \text{KCy} + \text{H}_2\text{O}$. Hence, to distinguish between cyanogen and hydrogen cyanide, pass the gas into a solution of potash-lye, and test for cyanates in the liquid by slightly acidifying with acetic acid, and adding a little solution of a cobalt salt. A bright blue coloration indicates the cyanates. The silver salts, etc., also have many analogous properties. If acetylene carboxylicdiamine, $\text{CO.NH}_2-\text{C}\equiv\text{C}-\text{CO.NH}_2$, be treated with a dehydrating agent, two molecules of water are removed, and a compound, $\text{N}\equiv\text{C}-\text{C}\equiv\text{C}-\text{C}\equiv\text{N}$, which can be regarded as carbon cyanide, C_2Cy_2 , remains in the form of white crystalline needles, melting at 20.5° to 21.0° and boiling at 76° (753 mm.). It readily burns in air, and has an irritating odour.

§ 4. Cyanic Acid and the Cyanates.

Potassium cyanate, $\text{KO}-\text{CN}$.—This salt is produced when potassium cyanide is slowly oxidized in air; and it is usually made by heating potassium cyanide or ferrocyanide with an oxidizing agent—litharge, red lead, potassium permanganate, etc.: $\text{KCy} + \text{PbO} = \text{Pb} + \text{KOCy}$. The potassium cyanate is extracted by lixiviating the mass with dilute alcohol, and concentrating the alcoholic solution by evaporation. Potassium cyanate is a colourless crystalline powder readily soluble in water and in dilute alcohol. The aqueous solution readily decomposes: $\text{KOCy} +$

$2\text{H}_2\text{O} = \text{NH}_3 + \text{KHCO}_3$. The corresponding acid, HOCy , decomposes so rapidly into carbon dioxide and ammonia: $\text{HOCy} + \text{H}_2\text{O} = \text{CO}_2 + \text{NH}_3$, that cyanic acid cannot be prepared by the decomposition of its salts with mineral acids.

Cyanic acid, HOCy .—If cyanuric chloride, Cy_3Cl_3 , be treated with water, cyanuric acid, $\text{H}_3\text{Cy}_3\text{O}_3$, is formed: $\text{Cy}_3\text{Cl}_3 + 3\text{H}_2\text{O} = 3\text{HCl} + \text{H}_3\text{Cy}_3\text{O}_3$, as a crystalline tribasic acid with a probable graphic formula:



If cyanuric acid be heated in a tube, and the vapours passed through a U-tube cooled by immersion in a freezing mixture, an unstable, liquid, monobasic acid with the empirical formula HCyO is obtained: $\text{H}_3\text{Cy}_3\text{O}_3 = 3\text{HCyO}$. The monobasic cyanic acid, HOCy , is the cyanogen analogue of hypochlorous acid, HOCl . If the temperature be raised above 0° , cyanic acid rapidly polymerizes into a hard, white, opaque mass called *camelide*, $(\text{HCyO})_n$, where n is not known.

Ammonium cyanate, NH_4CyO .—This salt is formed as a white crystalline powder when dry ethereal solutions of ammonia and cyanic acid are mixed together. It is also formed when a mixture of carbon monoxide and ammonia is passed over heated platinized asbestos, or subjected to the silent or spark electric discharge. On evaporating an aqueous solution of ammonium cyanate, F. Wöhler, in 1824, found that it underwent an intramolecular change, forming urea, $\text{CO}(\text{NH}_2)_2$ (p. 840), isomeric with ammonium cyanate. With potassium hydroxide, ammonium cyanate forms potassium cyanate.

§ 5. Thiocyanic Acid and the Thiocyanates.

If the alkaline cyanides be fused with sulphur, a change, analogous with the oxidation of the alkaline cyanide, occurs, and the so-called thiocyanates, or "sulpho-cyanides," are formed: $\text{KCy} + \text{S} = \text{KCyS}$. The fused mass, when cold, is lixiviated with dilute alcohol, and the alcoholic solution, when concentrated by evaporation, furnishes colourless deliquescent crystals of potassium thiocyanate, KCNs . Ammonium thiocyanate is conveniently made by digesting concentrated ammonia with carbon disulphide: $4\text{NH}_3 + \text{CS}_2 = \text{NH}_4\text{SCy} + (\text{NH}_4)_2\text{S}$. Ammonium thiocyanate undergoes a similar molecular change to that which occurs when ammonium cyanate is heated. At 140° , for instance, ammonium thiocyanate, NH_4SCN , passes into thiourea, $\text{CS}(\text{NH}_2)_2$, which is analogous with urea, $\text{CO}(\text{NH}_2)_2$. The thiocyanates give a blood-red ferric thiocyanate with ferric salts, and no coloration occurs with the ferrous salts if ferric salts be absent. Hence J. J. Berzelius' term *rhodanates*—from *rhodos*, a rose—for the thiocyanates. According to the ionic hypothesis, the red colour is supposed to be due to the un-ionized molecules, $\text{Fe}(\text{CyS})_3$, since neither the ferric ion Fe^{+++} nor the thiocyanates ion CyS^- are coloured. The red coloration is intensified if more ferric salt, or more thiocyanate be added to the solution, because the "addition of a common ion" causes part of the ionized salt to recombine to form molecules of the coloured ferric thiocyanate. When silver nitrate is added to a solution of potassium thiocyanate, a white flocculent precipitate of silver thiocyanate, AgCyS ,

separates. This is insoluble in dilute mineral acids. This reaction is the basis of Volhard's volumetric process for the determination of silver. Mercuric thiocyanate, $\text{Hg}(\text{CyS})_2$, made by adding mercuric chloride to a solution of potassium thiocyanate, is an insoluble powder which when washed and dried takes fire on ignition and forms a voluminous ash. Pellets made from the dry powder, when ignited, form long snake-like tubes—the so-called "Pharaoh's serpents." Thiocyanates are used for dyeing. A certain amount of ammonium thiocyanate is found in the "gas liquor," and in the "spent oxide" of the gas works.

Thiocyanic acid, HCyS .—This acid is made by distilling potassium thiocyanate with dilute sulphuric acid, or an intimate mixture of potassium hydrosulphate and potassium or barium thiocyanate, under reduced pressure, and passing the vapour through a tube containing calcium chloride to remove the vapour of water, and then through a U-tube cooled by a freezing mixture. At room temperatures, it is a colourless gas with a pungent odour; at -40° , a white crystalline solid is formed. The volatile liquid is quickly polymerized to a yellow mass of solid thiocyanic acid if removed from the freezing mixture, and this especially if concentrated sulphuric acid be present, for mineral acids accelerate the change. When thiocyanic acid is warmed with dilute sulphuric acid hydrolysis occurs and carbonyl sulphide, COS , is formed: $\text{HCyS} + \text{H}_2\text{O} = \text{NH}_3 + \text{COS}$; whereas cyanic acid under similar condition gives carbonyl oxide, that is, carbon dioxide. Thiocyanic acid was prepared by Winterl in 1790. It gives a white precipitate with silver nitrate; if traces of solid thiocyanic acid be present, the precipitate will be more or less yellow.

§ 6. Oxidation and Reduction.

This is a convenient place to recapitulate the meaning of the above terms. The word "oxidation" connotes the process of combination of oxygen with an element or compound; and "reduction," the reverse operation, namely, the withdrawal of oxygen from an oxy-compound. Loss of oxygen by heat, as in the "reduction" of mercuric oxide, p. 156, although it results in the removal of the oxygen from mercury, is not usually called reduction. The ideas associated with oxidation and reduction have been extended to include elements other than oxygen. For instance, the transformation of mercuric chloride, HgCl_2 , to mercurous chloride, HgCl , and finally to mercury by the action of stannous chloride, SnCl_2 , are processes of reduction. The stannous chloride is at the same time oxidized to stannic chloride, SnCl_4 , thus: $\text{SnCl}_2 + 2\text{HgCl}_2 = \text{SnCl}_4 + 2\text{HgCl}$. The two operations—oxidation and reduction—are reciprocal in that the oxidizing agent is reduced, and the reducing agent is oxidized by the process. In general, powerful oxidizing agents are readily reduced, and powerful reducing agents are readily oxidized. Similarly, the removal or addition of hydrogen is styled a process of oxidation or of reduction respectively, thus, acetylene, C_2H_2 , is reduced to ethylene, C_2H_4 , and ethylene is reduced to ethane, C_2H_6 , by nascent hydrogen. The change of a ferro- to ferri-cyanide is an oxidation process because it corresponds with a change of FeCy_2 to FeCy_3 , analogous with the transformation of FeCl_2 to FeCl_3 .

Oxidation is usually attended by an increase in the active valency and reduction by a decrease in the active valency of the central atom; or, as the ionic hypothesis would express it, the number of electrical charges on the ion is usually increased during oxidation, and diminished during reduction. Thus, when a solution of ferric chloride is reduced to ferrous chloride by treatment with stannous chloride, Sn^{++} becomes Sn^{+++} , and Fe^{+++} becomes Fe^{++} : $2\text{FeCl}_3 + \text{SnCl}_2 = \text{SnCl}_4 + 2\text{FeCl}_2$, two positive charges are gained by each atom of tin, and two are lost by each pair of iron atoms. When barium oxide changes to barium peroxide there is probably no change in valency. With the change $\text{H}_2\text{SO}_3 + \text{HOCl} = \text{H}_2\text{SO}_4 + \text{HCl}$, it is possible that the atom of sulphur passes from the quadri- to the sexi-valent state.

It is usual to say that oxidation is a process which involves the passage of a compound from a lower to a higher state of oxidation, by the addition of oxygen or of an acidic (electronegative) atom or radicle; or by the removal of hydrogen or an equivalent basic (electro-positive) atom or radicle. Reduction is the converse of the process of oxidation. An oxidizing agent is a substance which can engender oxidation as just defined; and conversely with a reducing agent.

Among the available oxidizing agents are: oxygen, ozone, the peroxides, and the higher oxides as well as the unstable basic oxides of silver, gold, etc.; the oxy-acids (nitric, nitrous, chromic, chloric, and the other oxy-acids of the halogens) and their salts; the halogens (chlorine, bromine, iodine); permanganic acid and its salts; potassium ferri-cyanide, etc.

Among the available reducing agents or deoxidizers are: hydrogen, unstable hydrides (hydrogen sulphide, hydrogen iodide, phosphine, arsine, stibine, etc.); carbon, carbon monoxide, sulphur dioxide, and the sulphites; phosphorous acid and the phosphites; hypophosphorous acid and the hypophosphites; potassium cyanide; potassium formate; ferrous, stannous, and chromous salts; the metals sodium, potassium, magnesium, aluminium, etc. These oxidizing and reducing agents have been discussed individually in earlier chapters.

Questions.

1. What is the action of concentrated sulphuric acid on any five of the following substances: (a) potassium nitrate, (b) sodium carbonate, (c) charcoal, (d) potassium cyanide, (e) potassium oxalate, (f) potassium iodide, (g) copper?—*St. Andrews Univ.*

2. How may potassium ferrocyanide be made? How are the following made from it: potassium cyanide, carbon monoxide, potassium ferri-cyanide, Turnbull's Blue? Write all the equations.—*Univ. Pennsylvania, U.S.A.*

3. Discuss any three of the following: (a) The action of heat on ammonium chloride; (b) The displacement of zinc by copper from a solution of a zinc salt containing excess of potassium cyanide; (c) The alkalinity of aqueous sodium carbonate solution; (d) The formula of ozone.—*St. Andrews Univ.*

4. 20 c.c. of a solution of hydrocyanic acid mixed with excess of potash require 50 c.c. of decinormal solution of silver nitrate to produce faint turbidity. What per cent. in the liquid?—*New Zealand Univ.*

5. Calculate the heat of formation of hydrogen cyanide, given: $\text{C} + \text{O}_2 = \text{CO}_2 + 96.9$ cal.; $\text{H}_2 + \text{O} = \text{H}_2\text{O} + 68.4$ cal.; $2\text{HCN} + 5\text{O} = 2\text{CO}_2 + \text{H}_2\text{O} + \text{N}_2 + 319.6$ cal.—*French Coll.*

6. Name two oxidizing agents and two reducing agents, and explain how they may be used in connection with the salts of iron. Under what conditions does the interaction occur in each case?—*Board of Educ.*

7. Outline methods by which the following cuprous compounds can be obtained from copper sulphate: (a) cuprous oxide, (b) cuprous chloride, (c) cuprous hydride, (d) cuprous cyanide, and describe briefly the appearance and properties of each. For what purposes is the chloride used in laboratories? Give reasons for the formula, CuCl or Cu_2Cl_2 , which you assign to this substance.

—*Board of Educ.*

8. Explain clearly the following phenomena in terms of the modern theory of solutions: (a) A solution of sodium carbonate is alkaline and a solution of sodium hydrogen carbonate is nearly neutral; (b) ferrous ammonium sulphate gives the usual reactions for iron, whereas potassium ferrocyanide does not.—*Bombay Univ.*

9. How is potassium ferrocyanide manufactured? How could you prepare from potassium ferrocyanide: (a) a dilute solution of prussic acid; (b) urea; (c) carbon monoxide?—*Calcutta Univ.*

10. If the atomic volume of oxygen when united to one element is 12.2, and to two elements 7.8; of sulphur united to one element 23.6, and to different elements 22.6; $\text{H} = 5.5$; $\text{C} = 11$; $\text{P} = 25.3$; and $\text{Cl} = 22.7$; calculate the molecular volume of phosphoryl chloride, and indicate which of the formulae $\text{P}(\text{OCl})\text{Cl}_2$ or POCl_3 is in agreement with the observed values 101.3. Also compare the calculated molecular volumes of sulphuryl chloride, SO_2Cl_2 , chlorosulphonic acid, HOClSO_3 , and carbon disulphide, CS_2 , when the observed values are respectively 86.3, 75.0, and 62.1.

11. Compare the molecular volumes of the following carbonates: calcite (sp. gr. 2.72); magnesite (sp. gr. 3.00); smithsonite (sp. gr. 4.42); siderite (sp. gr. 3.80).

12. What is meant by the term "hydrolysis"? Explain clearly the various factors which influence the degree of hydrolytic dissociation. State and explain what occurs when potassium cyanide, ferric chloride, and borax are separately dissolved in water; and what happens when ammonium carbonate is added to a solution of aluminium sulphate.—*Board of Educ.*

13. What information is conveyed by the use of terminals *-ous* and *-ic*? Illustrate your answer by examples, and state generally how *-ous* compounds are converted into *-ic* compounds, and *vice versa*.—*London Univ.*

14. What reactions occur when atmospheric nitrogen is passed over: (a) calcium, (b) calcium carbide, (c) barium carbide, (d) calcium hydride? What action, if any, has water on (1) the substances named, (2) the products obtained?—*Board of Educ.*

CHAPTER XL

SILICON

§ 1. Silica, or Silicon Dioxide.

NATURE furnishes some beautiful specimens of silica, and although the crystals are so attractive and so abundant, silica was for a long time regarded as an element, for it was argued that it is "la plus homogeneous, la plus simple, et la plus élémentaire des terres." The belief was also favoured by the infusibility and the resistance silica offers to attack by chemical reagents generally.

Silica is one of the most important compounds in the "half-mile crust" of the earth. It occurs abundantly in the mineral kingdom, and it is also common in the connective tissue of animals, fibres of vegetables, etc. The so-called *kieselguhr* or "diatomaceous earth" is a friable powder resembling chalk or clay. It is virtually a mass of siliceous skeletons of dead diatoms. Diatoms are minute plants with a siliceous casing or skeleton. The different varieties vary in size and shape but the casing usually tapers towards both ends. The organism lives as a scum on the surface of water—either floating free or attached to foreign objects. The organisms die, and sink to the bottom. The siliceous shells accumulate along with some decomposed vegetable matter. The resulting diatomaceous earths are not uncommon in peat-bogs and marshes on the margin of existing lakes. Some of the older deposits are consolidated—*e.g.* tripoli, so named from its occurrence in Oran (Tripoli). Silica, SiO_2 , occurs in nature free and combined with various bases to form numerous mineral silicates. Free silica occurs crystalline and amorphous. There are three main types of crystalline silica: quartz, tridymite, and cristobalite.

1. Quartz.—Quartz occurs in hexagonal prisms (trigonal system) terminating in hexagonal pyramids. A single quartz crystal weighing very nearly one ton has been reported from Calaveras (U.S.A.). The purest varieties of quartz—called *rock crystal*—are colourless; they have a specific gravity 2.67, and are hard enough to cut glass. The crystals are sometimes coloured with traces of various oxides. Thus, manganese oxide gives *amethyst quartz*; *smoky quartz* probably owes its colour to the presence of carbonaceous matter; *milky quartz* owes its opacity to the presence of innumerable air bubbles. Quartz also occurs massive in quartzite and quartzose rocks. Quartzose sands and sandstones are also more or less impure quartz. Rock crystal was one of the first crystallized minerals to attract the attention of the early philosophers, and they believed rock crystal to be a form of ice "so hard and dry that it becomes crystal"

—A. Magnus (1250); but, added Pliny, it is not easy to explain why the crystals grow six-sided. The Greek word for ice, *κρύσταλλος* (crystallos), has been extended to cover the whole science of crystals—crystallography. Agricola (1550) reported his belief that “rock crystal is not ice, but a denser product of cold.”

Crystals of quartz are sometimes found with hemihedral¹ faces inclining to the right in some specimens, and to the left in others, and, as L. Pasteur pointed out in 1860, “quartz crystals likewise separate themselves into two sets in relation to their optical properties, for one set deviates the plane of polarized light to the right, and the other set to the left, according to the same laws.” The former may be called right-handed, and the latter left-handed quartz. As in the case of tartaric acid, indicated on p. 628, the two crystals are enantiomorphous. The enantiomorphism of these crystals is illustrated by Fig. 326.

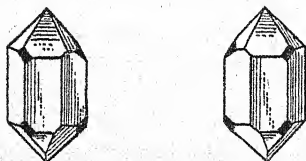


FIG. 326.—Enantiomorphous Quartz Crystals.

Silica dissolves in many molten silicates, and the effect on the freezing point of molten lithium silicate gives numbers corresponding with a molecular weight 120.8, meaning that the silica molecule in that solution is probably $(\text{SiO}_2)_2$. There are two forms of quartz with the transition point indicated in Fig. 327.

2. **Tridymite.**—This is a second variety of crystalline silica which was discovered by G. von Rath in 1868 in some andesitic rocks from the San Cristobel mountains at Pachuca (Mexico). Tridymite crystallizes in six-sided tabular crystals belonging to the triclinic system. Tridymite has been found in some meteorites. It is formed when quartz is heated for a long time at about 1000°. Hence tridymite is very common in silica bricks, etc., which have been heated in industrial furnaces. The transition temperature is about 870°. The transformation of tridymite back to quartz, below 800°, is exceedingly slow. The velocity of the change is accelerated in the presence of chlorides of the alkali metals, sodium tungstate, etc. The specific gravity of tridymite is 2.33 as contrasted with 2.67 for quartz. There are three forms of tridymite with the transition points 117° and 163° indicated in Fig. 327.

3. **Cristobalite.**—This is a third variety of crystalline silica which occurs in small octahedral (tetragonal) crystals up to about 2 mm. in size. It was discovered in the above-mentioned rocks at Pachuca. The specific gravity of cristobalite is practically the same as tridymite, being nearly 2.34. Cristobalite crystals are formed in silica bricks which have been heated some time to a temperature at which quartz begins to sinter.

¹ In normally formed crystals, every face has a similar face opposite to it, and not only opposite, but in every position where a face is possible, yet similarly situated with respect to similar axes—law of crystal symmetry. When a crystal shows all the faces required by the law of symmetry the crystal is said to be holosymmetrical or *holohedral*—from the Greek *ὅλος* (holos), whole; *ἑδρα* (hedra), base or face. A crystal less symmetrical may be derived from a holohedral crystal by the suppression of half its faces, when it is termed *hemihedral* from the Greek *ἡμι* (hemi), half—the tetrahedron, for instance, is the hemihedral form of the octahedron; or by the suppression of three-quarters of its faces, when it is termed *tetartohedral*—from the Greek *τέταρτος* (tetartos), a quarter.

There are two forms of cristobalite with the transition point at 270° as indicated in Fig. 327.

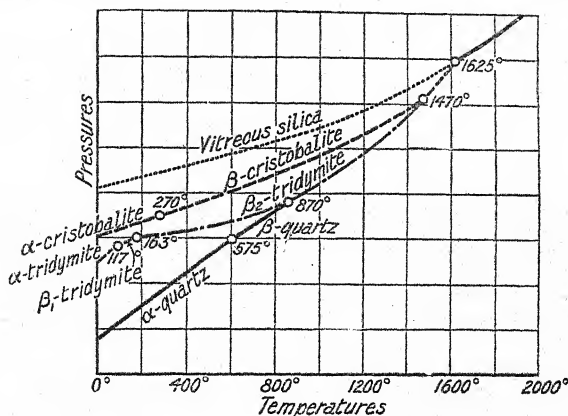


FIG. 327.—The Effect of Heat on Quartz.—After C. M. Fenner.

Amorphous silica.—Amorphous silica occurs in nature associated with 3 to 12 per cent. of water in the mineral *opal*, which may be colourless or tinted yellowish-brown, etc., with iron oxide, organic matter, etc. Chert, flint, chalcedony, and jasper contain more or less amorphous silica associated with quartz. The crystalline characteristics are so difficult to recognize that these minerals were once thought to be amorphous silica, and they are said to be cryptocrystalline—from the Greek *κρυπτός* (*cryptos*), hidden. The silica of diatomaceous earth is opaline and soluble in alkaline carbonate solutions, although the silica in some of the older deposits has begun to crystallize as quartz.

Properties of silica.—Silica melts to a colourless glass—quartz glass—in the oxyhydrogen blowpipe. The melting point of quartz is not well defined. Melting commences about 1600° . Silica can be vaporized in the electric furnace. The specific gravity of vitreous silica is about 2.22. The coefficient of thermal expansion of vitreous quartz is remarkably small—nearly 0.0000005—so that quartz glass can be very rapidly cooled without cracking. For instance, quartz glass can be heated red hot in the blowpipe and plunged in cold water without fracture; under the same conditions, ordinary glass—with a coefficient of thermal expansion of 0.000008—would shatter into small fragments. When heated for some time at about 1250° , the vitreous quartz passes into the crystalline condition (tridymite), and it will not then bear the sudden heating and cooling so well.

Silica is reduced by carbon in the electric furnace and forms carborundum (*q.v.*); it is reduced by magnesium to amorphous silicon. Crystalline and vitreous silica appear to be insoluble in water and in all acids except hydrofluoric acid. Fused silica is readily attacked by phosphoric acid and by the alkalis. Crystalline silica is slowly attacked by aqueous solutions of alkaline hydroxides, and carbonates, but the amorphous

variety is rapidly attacked. Silica is also attacked by superheated water, and a small quantity may pass into solution. The necessary conditions seem to prevail in deep-seated cavities in the earth. The water rising to the surface is cooled, and the pressure reduced. Some of the dissolved silica is then deposited at the mouth of the spring as a thick jelly. This afterwards changes into a hard white porous mass called *geyserite*. The Great Geyser of Iceland, for instance, is surrounded by a large mound or hillock of silica with a funnel-like cavity from which the geyser discharges. Similar geysers occur in the Hot Springs of New Zealand, the geysers and hot springs of Yellowstone Park, U.S.A., etc. Some *petrifying springs* owe their power to the silicic acid they hold in solution. In many cases—e.g. the mineral springs at Yellowstone Park—the alkalinity of the water facilitates the solution of the silica. The alkaline silicates are decomposed by the carbon dioxide of the atmosphere and the silica is deposited as geyserite or “siliceous sinter” in the neighbourhood. In general, the decomposition of the silicates by exposure to the atmosphere, or the weathering of silicates, furnishes amorphous or opaline silica.

Although chemically inactive at ordinary temperatures, silica acts as a powerful acid anhydride at high temperatures, combining with the bases and many metallic oxides to form more or less fusible silicates. When silica is fused with sodium carbonate, the carbon dioxide is expelled with effervescence. The reaction $\text{Na}_2\text{CO}_3 + \text{SiO}_2 = \text{Na}_2\text{SiO}_3 + \text{CO}_2$ formally resembles the action of sulphur dioxide on the same salt: $\text{Na}_2\text{CO}_3 + \text{SO}_2 = \text{Na}_2\text{SO}_3 + \text{CO}_2$. The sodium silicate, when dissolved in water, is decomposed by sulphuric acid and silicic acid is set free: $\text{Na}_2\text{SiO}_3 + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + \text{H}_2\text{SiO}_3$. The two end-products can be separated by dialysis. Hence sodium silicate behaves like other salts, and SiO_2 is called *silicic anhydride* in the same sense that SO_2 is called *sulphurous anhydride*. The more fusible silicates—e.g. lead silicate—are used in making glasses and pottery glazes. Potassium and sodium silicates are soluble in water, and the aqueous solution is sold as water glass, and the solid as soluble glass. The powerful acid character of silica at high temperatures turns on the fact that most of the acid anhydrides— SO_3 , P_2O_5 , etc.—volatilize at much lower temperatures, and consequently, as soon as ever so little, say, sulphur trioxide is displaced, the volatile anhydride passes away and ceases to compete with the silica for the base. At lower temperatures, sulphur trioxide rapidly displaces silica from the bases when competing under equal conditions.

Uses of silica.—In 1839 M. A. Gaudin showed that quartz can be melted and then worked like glass, but the fact attracted no particular attention until comparatively recent years. Quartz glass is used for the manufacture of elastic threads to suspend the delicate parts of electrical instruments. It is made into tubes, flasks, dishes, etc. Sandstone and quartzite are used for building stones, grindstones, whetstones, etc. Sand or sandstone is ground with a little lime or binding clay and made into refractory bricks, “ganister bricks,” “silica bricks,” “Dinas bricks,” etc. Many varieties of quartz are shaped into ornaments and gems. Diatomaceous earth—also called *tripoli*, *diatomite*, *kieselguhr*, or (wrongly) *infusorial earth*, etc.—is used as a polishing powder, in the manufacture of cement, soluble glass, and heat-insulating bricks. Its absorbent properties are utilized in pharmacy, etc., for the preparations of solids saturated with

liquid, antiseptics for dry dressings for wounds, etc.; it is also used for absorbing bromine, sulphuric acid, nitroglycerol, etc. The nitroglycerol then forms *dynamite*.

The sulphur analogue of silica—silicon disulphide, SiS_2 —can be made by heating an intimate mixture of amorphous silicon and sulphur at 100° , and then projecting the mass, in small portions at a time, into a red-hot crucible. By sublimation under reduced pressure, long, white, silky needles are obtained which burn in air to silica and sulphur dioxide; and with water furnish hydrogen sulphide and silica. Silicon monosulphide, SiS , has also been made in two different forms: The one, orange-yellow in colour, hydrolyzes in water furnishing hydrogen sulphide and silicic acid: $2\text{SiS} + 3\text{H}_2\text{O} = 2\text{H}_2\text{S} + \text{H}_2\text{SiO}_3 + \text{Si}$; the other, black in colour, furnishes silicon and silicon disulphide when heated: $2\text{SiS} \rightarrow \text{Si} + \text{SiS}_2$. Stannous sulphide, SnS , decomposes similarly and furnishes stannic sulphide and tin.

§ 2. Silicic Acids.

If the soluble alkaline silicates be treated with acids, an amorphous gelatinous mass called "silicic acid" is obtained. This is appreciably soluble in water and in acids, and is readily dissolved by dilute solutions of the alkali hydroxides and carbonates. The jelly-like mass, when dried in air, retains about 16 per cent. of water; and at 100° , about 13 per cent. The mass is then practically insoluble in water and acids. By further heating more and more water is expelled, thus, at 200° , about 5½ per cent. of water remains; at 300° , about 3 per cent.; and eventually, at about 500° , anhydrous silica is formed. The dehydration curve showing the vapour pressures of "hydrates" of different composition shows no "breaks" as would probably be the case if definite hydrates were formed. These facts correspond with the great variations C. Frémy found in the degree of hydration of natural opals dried *in vacuo*. The percentage amount of water varied from 1·5 to 12·1 per cent. If "silicic acid" which has been heated to 200° be exposed to a moist atmosphere, water is again absorbed, and the vapour pressure of the "hydrate" is greater than the original "hydrate" of the same composition. This shows that the water is probably less firmly retained by the re-hydrated silicic acid. A dilute solution of sodium carbonate—say, 5 per cent.—dissolves all the above-mentioned silicic acids. The rate of solution and possibly also the solubility of "silicic acid" is smaller the higher the temperature at which the hydrate has been heated. Silica which has been calcined at 1000° is dissolved with extreme slowness. Native quartz is almost insoluble in 5 per cent. sodium carbonate, but if finely powdered, appreciable quantities are dissolved in a short time.

When a solution of water glass (sodium or potassium silicate, say, Na_2SiO_3) is acidified with hydrochloric acid, some of the silicic acid separates as a gelatinous mass (hydrogel) and some remains in solution (hydrosol). If the solution be sufficiently dilute, the silicic acid will all remain in solution along with the excess of hydrochloric acid, and the sodium chloride formed in the reaction: $\text{Na}_2\text{SiO}_3 + 2\text{HCl} \rightleftharpoons \text{H}_2\text{SiO}_3 + 2\text{NaCl}$. The hydrochloric acid and the sodium chloride can be separated from the silicic acid by dialysis (Fig. 118). To avoid the trouble of

changing the water, the dialyzer illustrated in Fig. 328 can be used in place of the simpler form, Fig. 118, used by Graham. In the improved apparatus a current of water is kept circulating about the outside of the dialyzing membrane. The dialyzing surface is also relatively great so that the operation is much quicker than before.

The dialyzed acid does not give a luminous beam with Tyndall's optical test, showing that the dissolved particles are not large enough to scatter light. The coagulation of the solution is heralded by a feeble opalescence in the track of the beam of light. The opalescence gradually increases in intensity. A clear solution containing 5 per cent. of colloidal silica can be obtained as a hydrosol, by dialysis. This solution can be concentrated by boiling in a flask until it contains about 14 per cent. of silicic acid. The solution so prepared gelatinizes, or assumes the hydrogel condition, on standing a few days. The passage of silicic acid from the sol to the gel condition is retarded by the presence of a little hydrochloric acid, or alkali hydroxide, and is accelerated by a little sodium carbonate. If the clear solution of silicic acid be allowed to evaporate *in vacuo* at about 15°, a clear transparent jelly is obtained which, when dried over sulphuric acid, has approximately the composition H_2SiO_3 , that is, $\text{SiO}_2 \cdot \text{H}_2\text{O}$, and it is called metasilicic acid. An acid of the same composition has been made by dehydrating gelatinous silicic acid with 90 to 95 per cent. of alcohol. An acid of approximately the composition $\text{SiO}_2 \cdot 2\text{H}_2\text{O}$, that is, H_4SiO_4 , and called orthosilicic acid, has been made by dehydrating gelatinous silicic acid with absolute ether, and drying the amorphous white powder between folds of filter paper; the same compound is said to be formed by the action of silicon tetrachloride on water. Orthosilicic acid loses water on exposure to the air.



FIG. 328.—Proskauer's Dialyzer.

Like sulphurous and carbonic acids, the silicic acids dissociate so readily into water and acid anhydride, SiO_2 , that there is some doubt about the existence of definite hydrates. Indeed, it is generally believed that the water is not "chemically combined" with the silica, and the so-called silicic acids are said to be adsorption compounds (p. 949). The real existence of the definite hydrates just indicated is thus open to question. The fact that the powders prepared by the processes just indicated have approximately a composition corresponding respectively with $\text{SiO}_2 \cdot \text{H}_2\text{O}$ and $\text{SiO}_2 \cdot 2\text{H}_2\text{O}$ is attributed to chance. Be this as it may, a large number of compounds of silica with the bases—silicates—are known, and a large number occur in nature as definite minerals; and many others have been prepared artificially. These salts can be referred to unknown silicic acids.

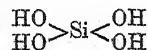
If a 1 per cent. aqueous solution of sodium silicate be decomposed by hydrochloric acid two kinds of silicic acid may be obtained. The freshly prepared acid— α -silicic acid—is not precipitated by a solution of egg-albumen, methylene blue, or sodium hydroxide; it changes slowly

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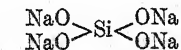
at ordinary temperatures—rapidly on warming— β -silicic acid which is soluble, but is precipitated by the reagents just mentioned. The β -acid is converted into the α -acid by warming its aqueous solution. Solutions of alkaline silicates of the type: R_4SiO_4 , R_2SiO_3 , $R_2Si_2O_5$, give solutions of the α -acid; and ordinary water glass, or the silicate, $Na_2Si_4O_9$, give the β -acid. Osmotic pressure phenomena indicate that the molecular weight of the β -acid approaches 49,000, and it is accordingly inferred that the molecule is very complex. The attempt to determine the molecular weights of colloidal suspensions by osmotic and related phenomena is irrational because the whole argument on which the methods are based is alone applicable to homogeneous media, and not to colloidal suspensions.

§ 3. The Silicates.

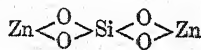
Starting with orthosilicic acid, H_4SiO_4 or $Si(OH)_4$, this is supposed to pass into metasilicic acid, H_2SiO_3 or $SiO(OH)_2$, by the loss of one molecule of water. These acids correspond respectively with the ortho- and metasilicates. Ethyl orthosilicate, $Si(OC_2H_5)_4$, is formed by the action of alcohol on silicofluoroform, $SiHF_3$. Among other orthosilicates, we have:



Orthosilicic acid.

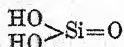


Sodium orthosilicate.

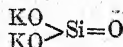


Zinc orthosilicate (willemite).

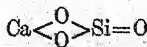
as well as *olivine*, Mg_2SiO_4 ; *zircon* (zirconium quadrivalent), $ZrSiO_4$; *tephroite*, Mn_2SiO_4 ; *fayalite*, Fe_2SiO_4 ; etc. Among the metasilicates:



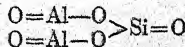
Metasilicic acid.



Potassium metasilicate.



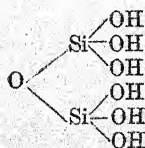
Calcium metasilicate (wollastonite).



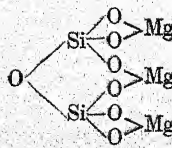
Aluminium metasilicate.

as well as *enstatite*, $MgSiO_3$; *rhodonite*, $MnSiO_3$; etc. *Andalusite*, *cyanite* or *disthene*, and *sillimanite* are three different crystalline forms of aluminium silicate, $Al_2O_3 \cdot SiO_2$, etc. *En passant*, it is interesting to note that a sodium metasulphosilicate, Na_2SiS_3 , has been made by melting sodium sulphide with silicon disulphide. The brownish-black mass gives silica and hydrogen sulphide when treated with water.

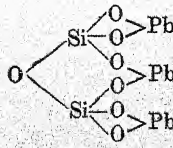
Two molecules of orthosilicic acid may be condensed into one molecule of orthodisilicic acid, $H_6Si_2O_7$, and hence we obtain a series of orthodisilicates:



Orthodisilicic acid.



Magnesium orthodisilicate (serpentine).



Lead orthodisilicate (barysilite).

Similarly, by the loss of one molecule of water between two molecules of metasilicic acid, the two molecules of the meta-acid can be condensed

to one molecule of metadisilicic acid, $\text{H}_2\text{Si}_2\text{O}_5$. Corresponding metadisilicates are well known. Similarly, ortho- and meta-trisilicic acids can be derived from three molecules of the respective acids. In addition to these silicates, hydrated, acid, and basic silicates are known. Many of the double silicates of aluminium and the bases are best referred to unknown aluminosilicic acids as indicated later. The system used in naming the silicates just indicated is conveniently summarized in Table LX.

TABLE LX.—NAMING THE SILICATES.

Name.	Hypothetical acid.	Silicate (R^1).			
		Mono-	Di-	Tri-	Poly-
Meta- . .	$\text{H}_2\text{O} \cdot \text{SiO}_2$	$\text{R}_2\text{O} \cdot \text{SiO}_2$	$\text{R}_2\text{O} \cdot 2\text{SiO}_2$	$\text{R}_2\text{O} \cdot 3\text{SiO}_2$	$\text{R}_2\text{O} \cdot n\text{SiO}_2$
Ortho- . .	$2\text{H}_2\text{O} \cdot \text{SiO}_2$	$2\text{R}_2\text{O} \cdot \text{SiO}_2$	$2\text{R}_2\text{O} \cdot 2\text{SiO}_2$	$2\text{R}_2\text{O} \cdot 3\text{SiO}_2$	$2\text{R}_2\text{O} \cdot n\text{SiO}_2$
Para- . .	$3\text{H}_2\text{O} \cdot \text{SiO}_2$	$3\text{R}_2\text{O} \cdot \text{SiO}_2$	$3\text{R}_2\text{O} \cdot 2\text{SiO}_2$	$3\text{R}_2\text{O} \cdot 3\text{SiO}_2$	$3\text{R}_2\text{O} \cdot n\text{SiO}_2$
...

Since we really know little more than the empirical formulæ of most of the silicates, the numerous attempts which have been made to classify the different silicates are more or less tentative, or speculative. The alkaline silicates are soluble in water, forming the so-called soluble glass; the other silicates are not usually soluble in water. Many of the simple silicates are attacked by hydrochloric acid, particularly if they have been roasted at a dull red heat. The silicic acid then separates as a gelatinous mass. The insoluble silicates are usually brought into solution for analysis by fusion with sodium carbonate, and the cold "cake" broken down by treatment with dilute hydrochloric acid. When the solution is evaporated to dryness, nearly all the silica separates in a form insoluble in dilute hydrochloric acid. The other constituents are dissolved by the dilute acid. The hydroxides of titanium, aluminium and iron are precipitated by ammonia from the filtrate from the silica. The iron and titanium can be separately determined in the precipitate and the alumina is obtained by difference. The magnesia and lime are determined in the filtrate from the precipitate by ammonia. The alkalis are determined on a separate sample.

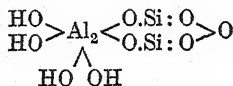
The formation of metallic silicates is well illustrated by a familiar experiment—*silica garden*: a litre beaker or bottle is filled with a solution of sodium silicate (sp. gr. 1.1) and crystals of, say, cobalt nitrate, cadmium nitrate, copper sulphate, ferrous sulphate, nickel sulphate, manganese sulphate, zinc sulphate, etc., are allowed to fall into the beaker so as to rest on different parts of the bottom. The whole is allowed to stand overnight in a quiet place, when plant-like shoots appear to grow from the crystals. The stalks and fronds have a form and colour characteristic of the metal salt used as "seed." Similar plant-like growths can be obtained by using other solutions—e.g. a granule of copper sulphate and sugar in a solution of potassium ferrocyanide and gelatine. It has been said that these growths mimic real plants. The growths are, of course, no more related to real plants than a living man is related to his statue. The phenomenon is an effect of osmotic pressure.

§ 4. The Aluminosilicates.

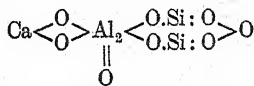
Certain compound silicates of aluminium with other bases appear to be "complex" aluminosilicates no more closely related to the silicates proper than the ferrocyanides are related to the cyanides. Just as it is convenient to refer the different silicates to more or less hypothetical silicic acids, so it is often convenient to refer many of the compound silicates to hypothetical aluminosilicic acids.

Alumino-monosilicic acid	$\text{Al}_2\text{O}_3 \cdot \text{SiO}_2 \cdot n\text{H}_2\text{O}$
Alumino-disilicic acid	$\text{Al}_2\text{O}_3 \cdot 2\text{SiO}_2 \cdot n\text{H}_2\text{O}$
Alumino-trisilicic acid	$\text{Al}_2\text{O}_3 \cdot 3\text{SiO}_2 \cdot n\text{H}_2\text{O}$
Alumino-tetrasilicic acid	$\text{Al}_2\text{O}_3 \cdot 4\text{SiO}_2 \cdot n\text{H}_2\text{O}$
Alumino-pentasilicic acid	$\text{Al}_2\text{O}_3 \cdot 5\text{SiO}_2 \cdot n\text{H}_2\text{O}$
Alumino-hexasilicic acid	$\text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2 \cdot n\text{H}_2\text{O}$

Thus, *kaolinite*, $\text{Al}_2\text{O}_3 \cdot 2\text{SiO}_2 \cdot 2\text{H}_2\text{O}$, appears to be an alumino-disilicic acid which has been called *kaolinic acid*, $\text{H}_4\text{Al}_2\text{Si}_2\text{O}_9$; *anorthite* or *lime-felspar*, $\text{CaO} \cdot \text{Al}_2\text{O}_3 \cdot 2\text{SiO}_2$, the calcium salt, and *soda-anorthite* or *carnegieite*, $\text{Na}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 2\text{SiO}_2$, the sodium salt of a similar acid, etc. Graphically, with the system indicated on p. 777, aluminium a triad, silicon a tetrad:

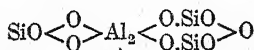


Kaolinite.

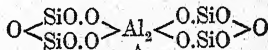


Anorthite.

With *natrolite*, $\text{Na}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 3\text{SiO}_2$, and *leucite*, $\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 4\text{SiO}_2$, we have salts of aluminotri- and aluminotetra-silicic acids respectively:

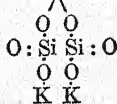
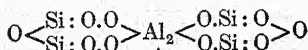


Natrolite.

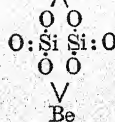
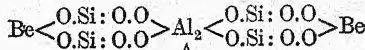


Leucite.

Potash-felspar, or *orthoclase*, $\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$, is regarded as the potassium salt, *albite* or *soda-felspar*, $\text{Na}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$, the sodium salt, and the *emerald* the beryllium salt, $3\text{BeO} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$, of an alumino-hexasilicic acid:



Potash felspar, or orthoclase.



Emerald.

There are many other systems of representing the constitution of the silicates. We know so little about the interior of the molecules that the value of systems of formulæ is mainly as a basis of classification. To avoid hypotheses, the silicate industries usually employ the old system of

J. J. Berzelius for the empirical formulæ, thus potash-felspar becomes $K_2O \cdot Al_2O_3 \cdot 6SiO_2$ and not $KAlSi_3O_8$; etc.

Felspathic rocks are fairly common, and thus a comparatively large supply of potash would be available for fertilizing purposes if a cheap enough method of extraction could be devised to enable the product to compete with that obtained elsewhere. Many methods have been devised although none as yet have proved commercially successful—*e.g.* J. G. A. Rhodin sintered a mixture of felspar, lime, common salt, and obtained the potassium as potassium chloride when the mass was lixiviated with water, and the solution crystallized. The leucite in Italy can be treated by several different methods: (1) When leached with sulphuric acid, it furnishes a solution of potassium aluminium sulphate which, on crystallization, yields alum, and the remaining silica dissolves in soda-lye yielding water-glass. (2) When leucite is treated with hydrochloric acid, potassium chloride and aluminium chloride are produced; and (3) when leucite is treated with caustic lime in an autoclave, caustic potash, alumina, and calcium silicate are formed.

Weathering of rocks.—Rocks generally show striking differences in behaviour when exposed for long periods of time; some remain hard and firm, others crumble to powder in a comparatively short time. Many ancient Egyptians and Grecian monuments show but slight symptoms of decay, whereas in other countries buildings made from apparently similar rocks soon deteriorate and are saved from disintegration only by a continued system of renovation. Calcareous building stones—*e.g.* the Houses of Parliament (London)—decay comparatively quickly when exposed to the acidic vapours which occur in the atmospheres of towns. When potash felspar and many other natural aluminosilicates are exposed to certain natural influences, they are finally converted into insoluble white crystalline or amorphous (colloidal) powder—clay, and other materials.

Formation of clays.—The early stages of the decomposition—weathering—of the felspathic rocks is indicated by the apparent clouding of the crystals of felspar; the felspar becomes more and more opaque; and finally disintegrates. As indicated above, the decomposition of silicate rocks exposed to weathering agents apparently furnishes colloidal silica—*e.g.* opal—and one of the last stages in the decomposition of felspar, and many other aluminosilicates, is clay. Consequently, the weathering of the aluminosilicates furnishes clay in a more or less colloidal condition. Granitic rocks, with felspar as a matrix, Fig. 3, disintegrate and leave behind the clay mixed with the more or less resistant varieties of mica, quartz, and other minerals which originally formed the granitic rock. The more important agents which facilitate the decomposition and disintegration—weathering—of the aluminium silicates are: (1) Volcanic gases (steam, hydrofluoric acid vapours, etc.). (2) Water draining from peat bogs, and coal beds. This water contains organic acids in solution. (3) Spring or rain water containing carbon dioxide, etc., in solution. The bases derived from the decomposition of the rocks are partly retained by the soil, and partly transformed into carbonates; the silicic acid is partly carried to the sea where it is utilized for building up the skeletons of various organisms, and partly retained on land where it unites with basic minerals forming steatite, serpentine, etc. The clay

may be leached by streams of water from the place where it was formed, transported from the hills, and deposited at lower levels. All kinds of *débris* from the rocks and soils, etc., over which the clays are carried may be transported along with the clay. Hence, the transported clays are the washings and sweepings of the hills which nature has accumulated as rubbish heaps in convenient places. These slow and insidious processes are continually at work levelling down the hills and mountains :

The hills are shadows and they flew
 From form to form and nothing stands ;
 They melt like mists the solid lands,
 Like clouds they shape themselves and go.—TENNYSON.

Transported clays are usually, but not always, less pure than the residual primary clays. The residual clays formed by the weathering of the less ferruginous granitic rocks, after an elaborate process of washing and settling, furnish white *china clay*, which has very nearly the empirical composition : $\text{Al}_2\text{O}_3 \cdot 2\text{SiO}_2 \cdot 2\text{H}_2\text{O}$. China clay is often called "kaolin," generally outside the industry. The disintegrated granite from which china clay is washed is called *china clay rock* ; a less disintegrated variety with more unweathered felspar is called *Cornish stone* ; and the grading of Cornish stone is based on the degree of disintegration or weathering it has undergone. The object of the washing china clay rock is to separate the china clay from the unweathered quartz, mica, etc. The crystalline mineral with the same ultimate composition as the purer forms of china clay is called *kaolinite*.

The term *clay* is applied industrially to a fine-grained mixture of various minerals which has these qualities : (1) It is plastic enough to be moulded when it is wet ; (2) It retains its shape when dried in spite of a certain amount of contraction ; and (3) When the moulded mass is heated to a high enough temperature it sinters together forming a hard coherent mass without losing its original contour. These properties have given clays a most important place—probably third or fourth—in the world's industries. Clays are used in the manufacture of building bricks, tiles, firebricks, crucibles, gas retorts, sanitary goods, pottery, etc. ; china clay is also extensively employed for filling paper, cotton, etc. *Fuller's earth* is a kind of non-plastic clay which disintegrates to a fine powder when placed in water. The washed powder was once used for "fulling" cloth—hence the name. It is now largely used for clarifying and deodorizing oils, etc. In this respect, it behaves like charcoal on coloured solutions.

Pottery and bricks.—British pottery is generally made from an intimate mixture of white-burning clay, with flint or quartz, and felspar or Cornish stone. The mixture is made into a slip (slurry) with water ; sieved ; and partly de-watered by filter press or other convenient method, so as to form a plastic mass. The mixture is moulded into the desired shape, dried, and fired between 1000° and 1200° according to the kind of ware being made. This forms the so-called "biscuit" body. A fusible mixture—containing lead boro-silicate, clay, felspar, etc., ground together to form a "slip" with water—is then spread over the surface of the "biscuit body," and the whole is refired to 900° or 1000° . The melted mixture covers the surface of the "body" with a glassy film or "glaze." There are many

modifications. The ware may be decorated by painting coloured oxides on the biscuit body before glazing; or by painting fusible enamels on the glaze and refiring; or the glaze itself may be coloured with suitable oxides. Glaze and body may be fired in one operation with or without a preliminary baking of the body. There are also considerable variations in the composition of the body and glaze. The chief varieties of pottery are *earthenware*—made from white-burning clays, Cornish stone, and flint; *vitreous ware*, and *ironstone*—made from a similar mixture but containing more fluxes and fired at a higher temperature; *hard porcelain*—made from clays, felspar, and quartz—with or without a little lime; *bone china*—made from bone ash, clay, and Cornish stone; *soft porcelain*—made from a calcareous marl and glassy frit; *parians*—made from a mixture of felspar and clay; and the commoner varieties of pottery—*terra-cotta ware*, etc.—made from special mixtures—often local clays glazed with a mixture containing galena, etc. There are all grades of porosity varying by insensible gradations from translucent glassy porcelain, to the most porous terra-cotta. No satisfactory basis of classification has been devised. *Drain pipes* are also made from local clays, which burn a buff or red colour, and glazed by throwing salt into the kiln. The salt decrepitates, volatilizes, and the vapours attack the surface of the clayware, covering it with a glass-like skin—*salt glaze*. *Tobacco pipes* (unglazed) are made from siliceous clays, that is, from clays containing more or less finely divided silica. *Firebricks* are made from refractory clays which soften at about 1580°. The refractory clay is moulded by hand or machinery, and fired to about 1100°–1200°. Common *building bricks* are usually made from less refractory clays fired at a low temperature.

Ultramarine.—Ultramarine or *lapis lazuli*—the sapphire of the Bible—occurs in nature as a blue, green, or violet coloured mixture of crystalline minerals, the most important of which is *lazurite* (not *lazulite*). It is considered to be a silicate of aluminium and sodium with some combined sulphur; but its constitution is by no means understood. Artificial ultramarine is a blue pigment made by calcining a mixture of china clay, sodium carbonate, charcoal, and sulphur in the absence of air. The green product is washed with water, dried, mixed with sulphur, and again roasted in air until the mass has acquired the required tint. Ultramarine is decomposed by acids with the evolution of hydrogen sulphide. It is used for neutralizing the yellowish tinge of sugar, cotton and linen goods, and in the laundry. It is also used as a blue pigment. The mineral has been almost superseded by the artificial product which is but one-fifth the price.

Imp **Glass.**—As previously indicated, glass is a solidified, undercooled solution of several silicates—most commonly potassium, calcium, and lead—and is made by fusing together a mixture of clean sand, limestone, or whiting or lime, sodium or potassium carbonate, and litharge or red lead in the right proportions. Traces of manganese dioxide or selenium are sometimes added to neutralize the yellow or green tinge due to the presence of ferrous or ferric oxide present as impurity in the ingredients used in making the glass. The mixture is melted in fireclay pots, and when the molten mass has cooled to the right temperature, a portion is collected at the end of an iron tube and brought to the desired shape by forcing it into a mould, or blowing into the tube and twisting or swinging the plastic

mass of glass as required. Details of the procedure vary with the particular objects being made. Rapidly cooled glass is brittle and liable to fracture, hence the glass is annealed in an annealing kiln where it can be cooled as slowly as desired. If cooled too slowly the glass devitrifies, *j.e.* crystallizes, p. 198.

It has been said that "few substances in daily use can be compared with glass in point of importance or utility. Glass is so common and so cheap that we are apt to lose sight of its unique qualities. It is durable, transparent, and easily cleaned; it can be readily coloured with metallic oxides; it is capable of taking a high polish; and, while in a fused condition, it can be made to assume almost any desired shape which it retains permanently when cold. It is an indispensable agent in most of the experimental sciences—chemistry, physics, astronomy, etc.—and it is difficult to imagine how many operations could be carried on without its aid. Without glass we should be more than centuries behind in astronomy, bacteriology, and biology." Window glass is a soda-lime silicate. This type of glass is sometimes called "soda-glass" or "soft glass," and it is used for making chemical glass ware. *Window glass, plate glass,* and glass for table ware, and bottles are also made from the same constituents in different proportions and of different degrees of purity. *Bohemian glass* is a potash-lime silicate. It is a hard glass and fuses only at a high temperature, hence it is used for making chemical apparatus designed to withstand high temperatures, gauge-tubes for boilers, etc. It also resists the solvent action of water better than soda-glass. *Jena glass* and *Bohemian glass* are varieties of potash-lime glass. *Flint glass* is a lead potash silicate. It is lustrous, and refracts light much better than other types of glass. It is used for making lenses for optical purposes. Some varieties are made into artificial gems and ornamental glass. *Cut glass* is a variety of lead glass which is ground or "cut" on emery or carborundum wheels. Beside these special admixtures, metallic oxides may be added to colour the glass. Translucent or white glass is made by the addition of bone ash, or fluorspar, or cryolite. Boric acid is also used in the manufacture of glass with a high refractive index for optical purposes.

§ 5. Carbon and Silicon Halides.

We shall see, later on, that the elements carbon, silicon, germanium, tin, and lead have a family relationship. They all form halides of the type CF_4 , CCl_4 , etc. The tetrachlorides, for instance, are all liquid at ordinary temperatures, and boil :

CCl_4	SiCl_4	GeCl_4	SnCl_4	PbCl_4
76°	59.6°	86°	113.9°	decomposes when heated.

Silicon tetrafluoride, or tetrafluomethane, SiF_4 .—This gaseous compound is important. It was discovered by K. W. Scheele in 1771. It was afterwards made by J. Priestley, and its composition determined by J. L. Gay-Lussac and J. Thénard, 1808; J. Davy, 1812; and J. J. Berzelius 1824. Silicon tetrafluoride is made by the direct action of fluorine on amorphous silicon. Carbon fluoride, it may be added, is made by the direct action of the elements. Silicon fluoride is also made by the action of hydrofluoric acid upon silica or on a silicate—*e.g.* glass: $\text{SiO}_2 + 4\text{HF}$

$\rightarrow 2\text{H}_2\text{O} + \text{SiF}_4$. The other fluorides of the family indicated above can be made by the action of hydrofluoric acid on the elements. Silicon tetra-

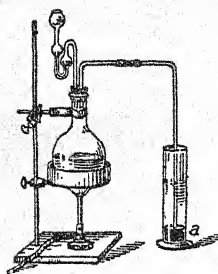


FIG. 329.—Preparation of Hydrofluosilicic Acid.

fluoride is usually made by the action of hydrofluoric acid derived from a mixture of calcium fluoride and sulphuric acid upon silica. The mixture is heated in a flask—illustrated in Fig. 329—fitted with a safety funnel containing mercury. An excess of sulphuric acid is used to absorb the water formed during the reaction.

Properties of silicon fluoride.—Silicon tetrafluoride is a colourless gas with a pungent odour resembling hydrogen chloride. The density of the gas is 104.2 (oxygen = 32). This corresponds with the formula SiF_4 . Silicon tetrafluoride condenses to a colourless liquid which boils at -65° under a pressure of nearly $2\frac{1}{2}$ atmospheres; and it solidifies when cooled still further, forming a white solid which melts at -77° . It can be sublimed without liquefaction at -90° at ordinary pressures. Glass is not attacked by dry silicon tetrafluoride. If the gas be passed over heated potassium it is decomposed with the separation of amorphous silicon: $\text{SiF}_4 + 4\text{K} = \text{Si} + 4\text{KF}$.

Hydrofluosilicic acid, H_2SiF_6 .—If silicon tetrafluoride be passed into water, it decomposes, gelatinous silicic acid is precipitated, and hydrogen fluoride is formed: $\text{SiF}_4 + 4\text{H}_2\text{O} = \text{Si}(\text{OH})_4 + 4\text{HF}$. The hydrogen fluoride immediately combines with a molecule of silicon tetrafluoride, producing an aqueous solution of hydrofluosilicic acid. The whole reaction is written: $3\text{SiF}_4 + 4\text{H}_2\text{O} = \text{Si}(\text{OH})_4 + 2\text{H}_2\text{SiF}_6$. In order to prevent the choking of the delivery tube by the separation of silicic acid when the silicon tetrafluoride is passed into water, it is well to let the delivery tube dip below a little mercury, *a*, Fig. 329, placed at the bottom of the vessel of water. The aqueous layer is frequently stirred to prevent the formation of channels of silicic acid through which the gas can escape into the atmosphere without coming in contact with the water. This is a good method of making hydrofluosilicic acid. The silicic acid is separated from the aqueous solution by filtration; the aqueous solution cannot be concentrated very much by evaporation because it decomposes into silicon tetrafluoride and hydrogen fluoride. A solid hydrate $\text{H}_2\text{SiF}_6 \cdot 2\text{H}_2\text{O}$ has been prepared.

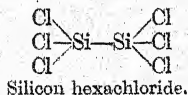
Hydrofluosilicic acid reddens blue litmus, and it is neutralized by the bases forming salts, fluosilicates. For instance, with potassium hydroxide, it forms potassium fluosilicate: $2\text{KOH} + \text{H}_2\text{SiF}_6 = \text{K}_2\text{SiF}_6 + 2\text{H}_2\text{O}$. Here the ion SiF_6^{--} behaves as a bi-valent anion. Most of the fluosilicates are fairly soluble in water, but the potassium and barium fluosilicates are dissolved with difficulty. Hence the use of hydrofluosilicic acid in testing for barium salts, and in the estimation of potassium. Hydrofluosilicic acid is used for giving wood a stone-like surface. The wood is first soaked in lime water and then treated with hydrofluosilicic acid. The acid is also used as an antiseptic in medicine.

Silicon tetrachloride, or tetrachlorosilicomethane, SiCl_4 .—We have seen that carbon tetrachloride, CCl_4 , is the final substitution product of methane, p. 850, or of carbon disulphide by chlorine. Germanium and

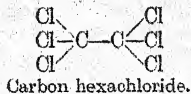
tin tetrachlorides— GeCl_4 and SnCl_4 —are made by the action of chlorine on the elements, and lead tetrachloride, PbCl_4 , by the action of chlorine on an hydrochloric acid solution of lead dichloride. Carbon tetrachloride cannot be made by the direct action of chlorine on carbon; but silicon tetrachloride can be made by heating silicon, ferrosilicon, silicon carbide, or an intimate mixture of carbon and silicon dioxide in a stream of chlorine: $\text{SiO}_2 + 2\text{C} + 2\text{Cl}_2 = \text{SiCl}_4 + 2\text{CO}$. The liquid which condenses can be freed from the excess of chlorine by shaking it with mercury, and redistilling. The colourless fuming liquid so obtained fumes in moist air. It has a vapour density and composition corresponding with SiCl_4 . It thus resembles carbon tetrachloride. Silicon tetrachloride boils at 53.3° and freezes at -89° ; carbon tetrachloride boils at 76° and freezes at -30° . Silicon tetrachloride is decomposed by water into silicic and hydrochloric acids: $\text{SiCl}_4 + 3\text{H}_2\text{O} = \text{H}_2\text{SiO}_3 + 4\text{HCl}$. Carbon tetrachloride is not acted upon by water, while the other chlorides of the family are decomposed in dilute aqueous solutions if hydrochloric acid be present. Silicon tetrachloride combines with ammonia, forming silicon tetramide, $\text{Si}(\text{NH}_2)_4$, and ammonium chloride, $\text{SiCl}_4 + 8\text{NH}_3 = 4\text{NH}_4\text{Cl} + \text{Si}(\text{NH}_2)_4$. This compound when heated forms silicon diimide, $\text{Si}(\text{NH})_2$, and silicon nitride, Si_3N_4 .

If the vapour of silicon tetrachloride be passed through a hot tube containing silicon, silicon hexachloride, Si_2Cl_6 , is formed: $3\text{SiCl}_4 + \text{Si} = 2\text{Si}_2\text{Cl}_6$. Silicon hexachloride is a mobile fuming liquid boiling between 146° and 148° , and freezing at -1° . Silicon hexachloride dissociates when heated to 350° : $2\text{Si}_2\text{Cl}_6 \rightleftharpoons \text{Si} + 3\text{SiCl}_4$, and the dissociation is practically complete at 800° . If, however, the silicon hexachloride be rapidly heated to 1000° , it has not time to dissociate to any great extent in passing through 350° to 1000° , and the compound is stable above 1000° . Allied phenomena have been previously studied, pp. 219, etc.

Perchlorethane, or carbon hexachloride, C_2Cl_6 , is analogous with silicon hexachloride; the former boils at 187° , the latter at 147° . The relations of the two are shown graphically:

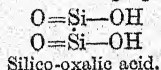


Silicon hexachloride.

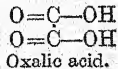


Carbon hexachloride.

Silicon hexachloride is hydrolyzed by water, forming silico-oxalic acid, $\text{Si}_2\text{H}_2\text{O}_4$, analogous with oxalic acid, $\text{C}_2\text{H}_2\text{O}_4$. The relation between these two acids is indicated graphically:

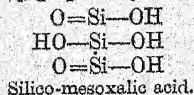


Silico-oxalic acid.

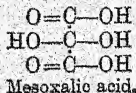


Oxalic acid.

The compounds Si_3Cl_8 and C_3Cl_8 are known; the former boils at 212° , the latter at 269° . Silicon octochloride is hydrolyzed by water, furnishing silico-mesoxalic acid, $\text{Si}_3\text{O}_2(\text{OH})_4$, analogous with mesoxalic acid, $\text{C}_3\text{O}_2(\text{OH})_4$, graphically:

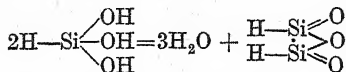


Silico-mesoxalic acid.



Mesoxalic acid.

Several other halogen compounds of silicon are known. When silicon is heated to redness in a current of dry hydrogen chloride, or when the gaseous products of the action of hydrochloric acid on copper silicide are cooled by liquid air, a compound—silico-chloroform, SiHCl_3 —is formed, boiling at 34° . The corresponding compound of carbon—chloroform, CHCl_3 —boils at 61° ; and the corresponding compound of germanium, GeHCl_3 , boils at 72° . Although chloroform is not hydrolyzed by water, silico-chloroform with water forms leucone, $\text{SiH}(\text{OH})_3$, also called silico-formic acid. The corresponding carbon compound is not known. Leucone is somewhat unstable, for it readily loses water, forming the compound, $\text{H}_2\text{Si}_2\text{O}_3$, which is called silico-formic anhydride:



When heated, silico-formic anhydride breaks down into silicon, hydrogen, and silica: $2\text{H}_2\text{Si}_2\text{O}_3 = \text{Si} + 2\text{H}_2 + 3\text{SiO}_2$. Probably this reaction means that silicon hydride, SiH_4 , and silica are first produced, and that the former decomposes at once into its elements. A compound silico-fluoroform, SiHF_3 , can be made by the action of silico-chloroform on stannic fluoride, SnF_4 . It boils at -80.2° , the corresponding carbon compound—fluoroform, CHF_3 —boils at 20° .

§ 6. Silicon.

History.—Silicon does not occur free in nature, but, as indicated in our study of silica, numerous oxygen compounds are known. The process of manufacturing glass from silicates has been known from ancient times, and J. J. Becher (1669) believed that these silicates contained a peculiar earth which he called *terra vitrescibilis* (vitrifiable earth); this is now called "silica." It was known in the 17th century that Becher's vitrifiable earth does not fuse when heated alone, and that a fusible glass is formed when it is heated with other earths. O. Tachenius (1660) noticed that the vitrifiable earth had acid rather than alkaline properties; K. W. Scheele (1773) showed it to have the characteristics of a refractory acid; and J. L. M. Smithson (1811) considered it to be an acid rather than an alkaline earth. J. J. Berzelius prepared amorphous silicon in 1823; and H. St. C. Deville prepared crystalline silicon in 1854.

Amorphous silicon.—This can be made by heating potassium or sodium in an atmosphere of silicon chloride or silicon fluoride: $\text{SiF}_4 + 4\text{K} = \text{Si} + 4\text{KF}$. The brown mass so formed is washed with water and hydrofluoric acid, heated at a dull red heat, and finally washed and dried. It is also formed by heating a mixture of sodium or potassium fluosilicate with metallic potassium: $\text{K}_2\text{SiF}_6 + 4\text{K} = \text{Si} + 6\text{KF}$. The brown mass is cleaned as before. Quartz is reduced to silicon when it is intimately mixed with magnesium powder and heated: $\text{SiO}_2 + 2\text{Mg} = \text{Si} + 2\text{MgO}$.

Amorphous silicon is a dark brown amorphous powder with a specific gravity 2.35. It melts at about 1500° , and volatilizes in the electric arc. When calcined in air, a surface skin of silica is formed which protects the element from complete oxidation. Silicon ignites in chlorine at about 450° , and burns to silicon tetrachloride. If silicon be heated with hydrogen

chloride, free hydrogen and silicon tetrachloride are formed: $\text{Si} + 4\text{HCl} = \text{SiCl}_4 + 2\text{H}_2$ (cf. p. 115). Silicon is insoluble in water and most acids; but it dissolves in hydrofluoric acid forming hydrofluosilicic acid: $\text{Si} + 6\text{HF} = 2\text{H}_2 + \text{H}_2\text{SiF}_6$. When boiled with alkaline hydroxides it forms hydrogen and alkaline silicate as indicated on p. 115.

Crystalline silicon.—Crystalline silicon is made by dissolving silicon in molten metals, and on cooling, part of the silicon separates from the solution in a crystalline condition. By passing a stream of silicon tetrachloride vapour over aluminium previously melted in an atmosphere of hydrogen, the volatile aluminium chloride passes on, and the silicon liberated by the reaction: $3\text{SiCl}_4 + 4\text{Al} = 3\text{Si} + 4\text{AlCl}_3$, dissolves in the molten aluminium. As the molten aluminium cools, silicon separates in long lustrous crystals. The aluminium can be separated by treatment with hydrochloric acid. Crystalline silicon is also made by heating a mixture of potassium or sodium silicofluoride, or powdered silica, with an excess of aluminium: $4\text{Al} + 3\text{K}_2\text{SiF}_6 = 3\text{Si} + 6\text{KF} + 4\text{AlF}_3$. The silicon dissolves in the excess of molten aluminium. The cold solution is treated with hydrochloric acid to remove aluminium, and with hydrofluoric acid to remove silica. Silica is reduced when heated with metallic magnesium: $\text{SiO}_2 + 2\text{Mg} = 2\text{MgO} + \text{Si}$; if an excess of magnesium be employed, magnesium silicide, Mg_2Si , is formed. Both it and magnesium oxide can be removed by treatment with hydrochloric acid. Crystalline silicon has been made commercially by heating quartz with coke in the electric furnace: $\text{SiO}_2 + 2\text{C} = 2\text{CO} + \text{Si}$. If too much coke be used, carborundum is formed. Silicon so prepared is sold in metallic-looking lumps and used in the manufacture of alloys.

Crystalline silicon forms dark grey opaque needle-like crystals or octahedral plates (cubic system). It is hard enough to scratch glass. Its specific gravity varies between 2.34 and 3, according to the temperature to which it has been heated. It burns when heated in chlorine and fires spontaneously in fluorine. Silicon is insoluble in acids, but dissolves in a mixture of nitric and hydrofluoric acids. It melts about 1500° , and distils in the electric furnace. Crystalline silicon slightly conducts electricity, amorphous silicon does not. Chemically, crystalline silicon resembles amorphous silicon, but it is not so active. Silicon combines with nitrogen, forming silicon nitride, and also with the metals, forming silicides. Siloxicon is the trade name for a greyish-green granular powder formed by heating a mixture of silica with carbon to about 2500° in an electric furnace. It varies in composition between $\text{Si}_2\text{C}_2\text{O}$ and $\text{Si}_7\text{C}_7\text{O}$. It is used as a refractory material when shaped into bricks, furnace linings, etc.

Silicon carbide, carbon silicide, carborundum.— SiC .—This compound is made by fusing a mixture of coke and sand in an electric resistance furnace—estimated temperature 3500° . The furnace is a large oblong box with permanent ends, and temporary sides. Large carbon electrodes are fitted into the two ends, and project into the furnace. Granulated coke is packed between the electrodes. A mixture of sand and coke, with some salt to make the mass fusible, and some sawdust to make the mass porous, is packed about the carbon core and held in place by the side walls of loosely packed bricks. The furnace is illustrated diagrammatically in Fig. 330. A powerful current of electricity is sent through the charge. The change which takes place is represented by the symbols:

$\text{SiO}_2 + 3\text{C} \rightarrow \text{SiC} + 2\text{CO}$. The operation is over in about eight hours. The furnace is then allowed to cool; the side walls are removed, and the

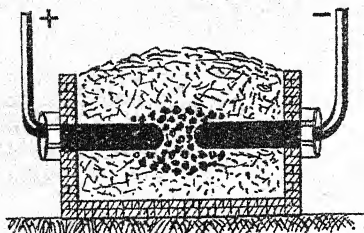


FIG. 330.—Carborundum Furnace (Diagrammatic).

silicon carbide removed. The best grades are found nearest the core. The product is crushed and treated with sulphuric acid to remove impurities; it is then washed, dried, and graded according to size.

Carborundum crystallizes in hexagonal plates when pure; it may be transparent and colourless, or vary in tint from emerald green to brown or black. The latter varieties are most common. The specific gravity is 3.2. It is not

attacked by acids—even hydrofluoric acid. It is decomposed by fusion with alkaline hydroxides. It is nearly as hard as the diamond, and accordingly is largely used as an abrasive powder and made into whetstones, hones, grinding wheels, polishing cloths, etc. It is also very refractory, and when mixed with clay has special uses as a refractory material for protecting furnace walls, etc.

Atomic weight.—The atomic weight determined by the analysis of silicon tetrachloride, silicon tetrabromide, etc., by different experimenters, lies between 27.95 and 28.38; the best representative value is taken to be 28.3, if oxygen be 16; and this number corresponds with the molecular weight deduced from the vapour densities of the volatile compounds of silicon by Avogadro's hypothesis. Dulong and Petit's rule does not apply so well unless the specific heat be taken at 300° , 0.2032.

§ 7. Hydrogen Compounds of Silicon.

There are six compounds of silicon and hydrogen with the general formula $\text{Si}_n\text{H}_{2n+2}$: silicon-methane or gaseous silicon hydride, SiH_4 , corresponding with methane, CH_4 ; silico-ethane or liquid silicon hydride, Si_2H_6 , corresponding with ethane, C_2H_6 , etc.; and two more or less doubtful products: silico-ethylene, Si_2H_4 , and silico-acetylene or solid silicon hydride, Si_2H_2 , corresponding with acetylene, C_2H_2 .

Silico-methane, SiH_4 , or silicane, or silane.—This gas is most conveniently made by the action of concentrated hydrochloric acid on magnesium silicide¹ whereby hydrogen gas containing 4 or 5 per cent. of silicane and a trace of silico-ethane is formed. The latter is spontaneously inflammable in air, the former is not. Hence the gas prepared by the above described process is spontaneously inflammable in air. This property can be illustrated by the method employed with phosphine, Fig. 238. The hydrochloric acid is placed in the flask, and the flask is filled with hydrogen gas; the current of hydrogen is shut off, and the magnesium silicide is dropped into the acid. The bubbles of gas ignite as they rise to the surface of the water, forming rings of silicon dioxide.

¹ Magnesium silicide, Mg_2Si , is made by heating sand with an excess of magnesium powder or by fusing together 40 parts by weight of anhydrous magnesium chloride with sodium fluosilicate, 35; sodium chloride, 10; and sodium, 20.

By passing the dried gases from the magnesium silicide and the acid through a tube surrounded by liquid air, both the SiH_4 and the Si_2H_6 are condensed to liquids. By fractional distillation of the condensed mass silico-methane is obtained as a colourless gas, not spontaneously inflammable at atmospheric pressure, but inflammable in air if slightly warmed or subjected to reduced pressure. Silico-methane burns with a bright flame, forming silica and water: $\text{SiH}_4 + 2\text{O}_2 = \text{SiO}_2 + 2\text{H}_2\text{O}$. When passed into alkaline solutions it decomposes, forming alkaline silicates and hydrogen: $\text{SiH}_4 + 2\text{KOH} + \text{H}_2\text{O} = \text{K}_2\text{SiO}_3 + 4\text{H}_2$. The boiling point is -112° , and the melting point -185° . When heated to 400° it decomposes into its elements, amorphous silicon and hydrogen: $\text{SiH}_4 = \text{Si} + 2\text{H}_2$. The specific gravity and boiling point of the liquid and the melting point of the solid are indicated in the following scheme along with the values for methane:

	Sp. gr. liquid.	Boils at	Melts at
Methane, CH_4	0.415 (at -164°)	-160°	-184°
Silane, SiH_4	0.680 (at -185°)	-112°	-185°

Silico-ethane, Si_2H_6 .—Hydrogen silicide remains as a liquid when the silico-methane has been distilled from the liquefied gases obtained by the action of hydrochloric acid on magnesium silicide, as described above. It is a colourless liquid, boiling at -15° , and solidifying at -132.5° . It can be heated to 100° , in the absence of air, without decomposition; but at 200° it decomposes into its elements hydrogen and amorphous silicon: $\text{Si}_2\text{H}_6 = 3\text{H}_2 + 2\text{Si}$. It is spontaneously inflammable in air, burning to water and silica. Like silico-methane, this compound reacts vigorously with free halogens. A comparison of the properties of silico-ethane with those of ethane and diborane shows that

	Sp. gr. liquid.	Boils at	Melts at
Ethane, C_2H_6	0.45	-89.3°	-171.4°
Diborane, B_2H_6	—	-87°	-169°
Silico-ethane, Si_2H_6	0.69	-15°	-132.5°

Similarly for the third member of the series, $\text{Si}_n\text{H}_{2n+2}$:

	Sp. gr. liquid.	Boils at	Melts at
Propane, C_3H_8	0.54	-45°	below -195°
Silico-propane, Si_3H_8	0.743	0.53°	-117.4°

Similarly for the fourth member of the series

	Sp. gr. liquid.	Boils at	Melts at
Butane, C_4H_{10}	0.60	0.6°	-135°
Tetraborane, B_4H_{10}	—	16°	-112°
Silico-butane, Si_4H_{10}	0.825	109°	-90°

Silico-pentane, Si_5H_{12} , and silico-hexane, Si_6H_{14} , have also been made. A silico-ethylene, Si_2H_4 , has been reported; but if it really exists very little is known about it. Similar remarks apply to the so-called silico-acetylene, $(\text{Si}_2\text{H}_2)_n$, which is supposed to be a solid formed when calcium silicide¹ is decomposed by hydrochloric acid: $\text{CaSi}_2 + 2\text{HCl} = \text{Si}_2\text{H}_2 + \text{CaCl}_2$. It is a yellow crystalline solid. The hydrides of silicon are not very stable. Like many of the other hydrides, they act as reducing agents. Thus

¹ Calcium silicide. Made by heating lime, silica, and carbon in an electric furnace.

silico-methane reduces silver nitrate with the separation of silver and silicon: $\text{SiH}_4 + 4\text{AgNO}_3 = 4\text{HNO}_3 + 4\text{Ag} + \text{Si}$. Copper sulphate is not so easily reduced as silver nitrate, and a compound of copper and silicon—copper silicide— Cu_2Si , is formed: $\text{SiH}_4 + 2\text{CuSO}_4 = \text{Cu}_2\text{Si} + 2\text{H}_2\text{SO}_4$. The more stable methane does not act like silico-methane. Acetylene, it will be remembered, forms acetylides or carbides with silver nitrate and copper sulphate. Many of the less stable hydrides—like silicon hydride, hydrogen sulphide, and hydrogen iodide—reduce in virtue of the hydrogen liberated during their decomposition; other hydrides reduce by the direct oxidation of their elements. Thus, with arsenic hydride and silver nitrate, as previously described, arsenious acid and silver are formed during reduction.

Questions.

1. Explain thoroughly how quartz may be put into solution and further treated so as to recover it as silica. What experimental evidence have we that arsenic is an amphoteric element?—*Princeton Univ., U.S.A.*

2. Describe the preparation, properties, and reactions of the compounds of silicon with hydrogen, with chlorine, and with fluorine.—*Aberdeen Univ.*

3. How can (a) silicon, (b) silicon carbide be obtained from sand, and for what purpose are these substances used? How can silicon chloride be obtained from silica, and converted into silica?—*Sheffield Univ.*

4. Compare and contrast the elements carbon and silicon by a discussion of their analogous inorganic compounds.—*Sheffield Univ.*

5. Explain the meaning of the term "anhydride." Describe the preparation of sulphur dioxide, chromium trioxide, and silicon dioxide, and the experiments by which you would prove each to be an anhydride.—*London Univ.*

6. What is the composition of felspar, caespar, fireclay, and fluorspar? From which of these and by what process could you produce silica?—*Board of Educ.*

7. Orthoclase felspar has the composition $\text{K}_2\text{O} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$. Explain the methods by which the following substances could be obtained from it in a state of purity: alumina, potassium chloride, potash alum, silica.—*Board of Educ.*

8. Given the heats of formation $2\text{Ca} + \text{O}_2 = 2\text{CaO} + 290 \text{ Cals.}$; $2\text{Si} + \text{O}_2 = 2\text{SiO}_2 + 382 \text{ Cals.}$; and $2\text{CaO} + 2\text{SiO}_2 = 2\text{CaSiO}_3 + 66.2 \text{ Cals.}$, show that the heat of formation of calcium metasilicate, CaSiO_3 , from its elements is 369.1 Cals.

9. Show that in the complete oxidation of the silicon carbide, Si_3C_4 , in a crucible, no loss in weight would be observed before and after calcination—neglecting any gaseous products which pass from the system and are not weighed.

10. How are the following substances prepared: Iodic anhydride, carbon oxysulphide, silicon tetrachloride, silicon chloroform, phosphorus pentafluoride, and chlorosulphonic acid?—*Science and Art Dept.*

11. How can you prove that the constituent of flint has properties similar to other acids, and how can this substance be obtained dissolved in water?—*London Univ.*

12. Silica has no taste and does not act on litmus. Why do we call this substance silica acid?—*London Univ.*

13. Make a comparison of the compounds of carbon and silicon, so as to reveal the relationship existing between them. What suggestion can be made to account for the fact that carbon dioxide is a gas, whereas silica is a solid of high melting point? What abnormalities are found in connection with the specific heats of the two elements?—*Board of Educ.*

14. How would you explain the following: (a) a solution of sodium silicate is readily decomposed by weak acids, but sodium sulphate heated at a high temperature with silica forms sodium silicate? (b) Barium peroxide heated at a certain temperature under reduced pressure forms barium monoxide with the liberation of oxygen, whereas barium monoxide heated at the same temperature under increased pressure combined with oxygen forms barium dioxide?—*Board of Educ.*

15. In what respects does silicon resemble carbon? Give the formulæ of the oxides of silicon. Explain what happens when a mixture of powder of flint and of sodium carbonate is raised to a bright red heat. How would you prepare pure silicic acid from the fused mass?—*London Univ.*

16. Give an account of the element silicon. Discuss the classification of the silicates.—*Cape Univ.*

17. Name minerals which consist of crystallized, of amorphous, and of mixtures of amorphous and crystallized silica respectively. Describe how you would prepare pure silica from flint.—*London Univ.*

18. Describe the changes which take place when quartz is heated to a high temperature.

19. The analysis of an amorphous silicon carbide furnished: Silicon 69.41 per cent.; 29.18 per cent. of carbon; 0.59 per cent. of iron and aluminium. What is the best representative formulæ of the substance?

20. Write a short essay on the chemical relationship between carbon and silicon.—*Science and Art Dept.*

21. Describe the preparation of the metal aluminium from clay, giving the chemical equations of the reactions.—*London Univ.*

CHAPTER XLI

TIN, LEAD, AND SOME RELATED ELEMENTS

§ 1. Germanium.

IN 1885 Weisbach discovered a silver mineral—*argyrodite*—in a mine at Freiberg (Saxony). C. Winkler analyzed the mineral, but found his analysis to be about 7 per cent. too low :

Silver.	Sulphur.	Ferrous oxide.	Zinc oxide.	Mercury.	Total.
74.72	17.13	0.66	0.22	0.31	93.04

Winkler traced the discrepancy to the presence of 6.93 per cent. of a new element, precipitated as sulphide in the "hydrogen sulphide group." This element he called germanium from the Latin name *Germania*. The new element proved to be bi- and quadri-valent. The quadrivalency of germanium was established by C. Winkler's study of germanium tetraethyl, $\text{Ge}(\text{C}_2\text{H}_5)_4$. The compounds corresponding with bivalent germanium resemble the compounds of silicon and carbon; and compounds corresponding with quadrivalent germanium resemble tin and titanium compounds. The analysis and vapour density of germanium tetrachloride correspond with an atomic weight 72.5 (oxygen = 16). This number agrees roughly with the atomic weight calculated from the specific heat 0.08 by Dulong and Petit's rule. Neglecting the small admixtures of iron, zinc, and mercury, the analysis of argyrodite thus corresponds with $3\text{Ag}_2\text{S}.\text{GeS}_2$.

2. Tin.

History.—Discoveries of tin in Egyptian tombs show that the metal was fairly common in olden times. It is not certain if the Hebrew word "bedil" in the Pentateuch, translated by the Greek word *κασσίτερος* (cassiteros), and by the Latin word *stannum*, really means tin. The word "stannum" appears to have been used by the Romans to designate certain alloys containing lead. It is not certain whether the Phœnicians obtained their tin from India, Britain, or Iberia. The resemblance between the Sanscrit word "castira" and the Greek "cassiteros" has been used as an argument in favour of the Indian origin of Phœnician tin. Pliny states that "cassiteron" was obtained from "Cassiterides (British Isles) in the Atlantic Ocean." This no doubt refers to the tin then obtained from the Cornish mines, for "certain islands north of Spain" were often referred to as the *insulæ cassiterides*—tin islands. The Romans appear to have distinguished lead from tin by calling lead "plumbum nigrum," and

tin "plumbum candidum." The word "stannum" was later restricted to tin proper. The alchemists called tin "Jupiter," and represented this metal by J , the symbol for the planet Jupiter.

Occurrence.—There are several reports of the occurrence of metallic tin in nature—Bohemia, Bolivia, New South Wales, Nigeria, etc. Practically, *tinestone* or *cassiterite* is the sole source of commercial tin. This mineral occurs in tetragonal crystals coloured brown or black by impurities, chiefly iron. Cassiterite is stannic oxide, SnO_2 , contaminated with more or less arsenical pyrites, copper pyrites, tungstates, and various metallic sulphides. *Lode tin* or *vein tin* is cassiterite which is obtained from veins or lodes in primary deposits; while "stream" tin is cassiterite from alluvial secondary deposits where it occurs in more or less rounded lumps. The miners speak of tinestone as "tin" or *black tin* to distinguish it from the metal which is called *white tin*. The complex sulphide ore, *stannite*, or *tin pyrites*, is a sulphide of copper, tin, iron, and sometimes zinc. It is comparatively rare. About one-third of the world's output of tin is produced in the Malay peninsula—e.g. Banca, which furnishes the so-called *Straits' tin*. Tin is also produced in the Malay archipelago, Bolivia, Nigeria, Australia, Cornwall, South Africa, Bohemia, and Saxony, etc.

Extraction.—The method of extracting tin from its ore, the metallurgy of tin, is comparatively simple, owing, no doubt, to the simplicity of the ore—black tin. The ore is first concentrated by washing away the earthy impurities. The high specific gravity of tinestone—6.8 to 7.0—enables this to be done without much trouble, as in the case of washing gold (*q.v.*). This process usually works well with stream tin; but vein tin usually requires more complex treatment. The crushed ore is first washed to remove earthy matters. The arsenic and sulphur are removed by an "oxidizing roast." The tungsten and the residue left after the calcination of the pyrites are removed by passing the calcined ore through the intense magnetic field of an "electro-magnetic separator."

The extraction of tin from the concentrates, as they are called, involves the reduction of the ore in a blast furnace or in a reverberatory furnace. To do this, the "concentrated" ore—tinestone—is heated with coal in, say, a reverberatory furnace. The oxide is reduced: $\text{SnO}_2 + 2\text{C} = 2\text{CO} + \text{Sn}$. The molten tin which collects on the bottom of the furnace is drawn off and cast into ingots or blocks—block tin—which contain about 99.5 per cent. of metallic tin. The slag obtained in this operation is also worked up to recover the 20 to 40 per cent. of metal it contains.

Refining tin.—Tin is refined by heating it, at a temperature as little as possible above the melting point of the metal, on the sloping hearth of a reverberatory furnace. The tin flows down the hearth and leaves the oxidized metallic impurities as "refinery dross" behind. This tin may be further purified by stirring the molten metal with a billet of wood. The metal is agitated by the bubbling of the rising gases, and this continually exposes fresh portions of the molten metal to the oxidizing action of the air. The impurities which collect on the surface as a "dross" are skimmed off. The refinery drosses containing a large percentage of tin are re-smelted with the ore.

Tin is recovered from scrap tin, tin plate, etc., by treating these materials with some solvent, e.g. chlorine. Electrolytic methods of extraction have not been very successful.

Properties.—Tin is a white lustrous metal with a pale blue tinge. The metal retains its lustre unimpaired by exposure to air. The metal is soft enough to be cut with a knife, but it is harder than lead, and not so hard as zinc. Tin is very ductile, for it can be beaten into foil—tinfoil—0.01 to 0.1 mm. thick, and drawn into wire. The ductility of tin is greatest at about 100°; at 200° the metal is brittle enough to be pulverized into powder. Tin shows a marked tendency to crystallize on solidification. If a bar of tin be bent, it emits a low crackling noise—"tin cry"—said to be due to the rubbing of the crystal faces upon one another. If the surface of a block of tin be treated with warm dilute aqua regia, the surface of the metal immediately assumes a crystalline appearance. The crystals are best shown by cooling molten tin in a crucible until part has solidified, and pouring out the liquid portion. The walls of the crucible will be lined with crystals of tin. Tin appears to be dimorphous, for electrolytic tin—i.e. tin deposited from a solution by the electrolysis of a tin salt—and malleable tin, form tetragonal crystals; while the "brittle" tin, mentioned above, is rhombic. Tetragonal tin passes into the rhombic form between 170° and 200°. Tin melts at 232°, and boils at about 2275°. A perceptible volatilization occurs at 1200°. The metal takes fire when heated between 1500° and 1600°, burning with a white flame to stannic oxide. When the metal is heated just above its melting point in air for some time, it forms a yellowish-white scum which is also stannic oxide.

Grey tin.—When cooled to a low temperature, tin crumbles to a grey friable powder. Several cases have been reported where tin, during an exceptionally cold winter, has crumbled to powder; for instance, A. L. Erdmann (1851) noticed some tin organ-pipes in the church at Zeitz crumbled to powder; and later, the tin buttons of some military uniforms were found to have crumbled to powder while in the dépôt during a cold winter. The disease is called the "tin pest." The afflicted tin first tarnishes, then shows faint radiating formations, and then wart-like formations, and finally the metal crumbles to a grey pulverulent mass. The afflicted metal contains two kinds of tin—the one, ordinary white bright tin; and the other grey, dull, pulverulent tin. Grey tin appears to be a third allotropic modification of the element.

	170°	18°
	Rhombic tin. \rightleftharpoons Tetragonal tin. \rightleftharpoons Grey tin.	
Specific gravity	7.25	6.55
Specific heat	0.0538	—
		0.0496

The transition temperature is 18°. Hence, excepting in warm weather, all ordinary white tin is in a metastable condition; but, as E. Cohen has pointed out, the speed of the transformation is very slow at ordinary temperatures. The transformation proceeds with a maximum velocity at -48°, especially if the tin be in contact with an alcoholic solution of "pink salt," $\text{SnCl}_4 \cdot 2\text{NH}_4\text{Cl}$. At lower temperatures, the velocity again slows down. The disease is contagious, for if a piece of tin which has already commenced to change be allowed to remain in contact with a piece of ordinary white tin, the unchanged tin is more quickly affected with the "disease." The specific heat of ordinary white tin is 0.0538, and of grey tin, 0.0496.

The action of acids and alkalis.—Tin dissolves slowly in dilute

hydrochloric acid, and fairly rapidly in the hot, concentrated acid, forming stannous chloride, SnCl_2 , and hydrogen: $\text{Sn} + 2\text{HCl} \rightarrow \text{SnCl}_2 + \text{H}_2$. If a little tin be placed in a test-tube with a little hydrochloric acid, there is little or no action, but if a drop of a solution of a platinum salt be added to the acid, a brisk reaction sets in; the platinum appears to act catalytically. Tin is but slowly attacked by cold sulphuric acid, but the hot concentrated acid dissolves the metal, forming stannous sulphate and sulphur dioxide: $\text{Sn} + 2\text{H}_2\text{SO}_4 \rightarrow \text{SnSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$. The action of nitric acid depends upon the temperature and concentration of the acid. With cold, very dilute nitric acid, stannous nitrate, $\text{Sn}(\text{NO}_3)_2$, is formed: $4\text{Sn} + 10\text{HNO}_3 = 4\text{Sn}(\text{NO}_3)_2 + 3\text{H}_2\text{O} + \text{NH}_4\text{NO}_3$, with possibly a little stannic nitrate, $\text{Sn}(\text{NO}_3)_4$. With nitric acid of moderate concentration (specific gravity 1.24), copious fumes are evolved and a bulky white precipitate of a hydrated oxide, meta-stannic acid, separates. Highly concentrated or pure nitric acid is practically without action on the metal. Boiling alkaline hydroxides form alkaline stannates, e.g. K_2SnO_3 , and hydrogen: $\text{Sn} + 2\text{KOH} + \text{H}_2\text{O} \rightarrow \text{K}_2\text{SnO}_3 + 2\text{H}_2$.

Atomic weight of tin.—The combining weight of tin has been determined by the analysis or synthesis of the oxide, chloride, bromide, sulphide, and the potassium and ammonium stannichlorides. The best available numbers show that this constant probably lies somewhere between 118.98 and 119.10. The best representative value is taken to be 119—oxygen = 16. The vapour density of the volatile compounds of tin; observations on the isomorphism of stannic and titanic oxides; and the specific heat, 0.055, by Dulong and Petit's rule, all point to this number, 119, as the atomic weight of tin.

Uses.—The resistance of tin to ordinary corrosive agents is utilized in protecting iron from rusting, "tin plate" is made by dipping thin sheets of steel into molten tin, whereby the steel is coated with a thin film of tin. The plated tin so made is used in the manufacture of tin cans, and similar articles—tintacks, for example, are tinned iron tacks, tin alone would be too soft to drive into wood. Copper coated with tin is also used for cooking vessels. Tin amalgam is used in coating mirrors.

Alloys.—Many useful alloys contain tin. The addition of tin to lead lowers the melting point of lead; and the addition of lead to tin lowers the melting point of the tin. Tin melts at 232° , and lead at 327° ; an alloy of 37 per cent. lead with 63 per cent. of tin melts at the eutectic temperature 180° , as indicated in Fig. 331. *Common solder* has one part of tin to one part of lead, but solders generally vary from 66 to 33 per cent. of tin. *Pewter* is a tin-lead alloy containing 75 per cent. of tin with 25 per cent. of lead. It will be observed

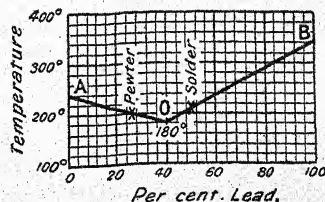


FIG. 331.—Melting Points of Tin-lead Alloys.

that an alloy of two metals or a salt solution may appear to have two freezing points: (1) the temperature at which an excess of one constituent freezes along the lines AO , OB , Fig. 331; and (2) the temperature at which the eutectic freezes *en bloc*. During the cooling of plumber's solder, for instance, say tin 40, lead 60, solid lead begins to separate at 240° ,

and continues separating until the mother liquid contains 37 per cent. of lead, when the whole mass solidifies. Between these two temperatures, 180° and 240°, the cooling alloy seems to be in a "pasty" condition, and this enables the plumber to "wipe" a joint being made with such an alloy.

§ 3. Stannous Oxide and Hydroxide.

When stannous chloride, SnCl_2 , is dissolved in an excess of water, a white precipitate of stannous oxychloride, $\text{Sn}(\text{OH})\text{Cl}$, is formed: $\text{SnCl}_2 + \text{H}_2\text{O} \rightleftharpoons \text{Sn}(\text{OH})\text{Cl} + \text{HCl}$. If stannous chloride be treated with an alkali, say sodium carbonate, or sodium hydroxide, stannous hydroxide, $\text{Sn}(\text{OH})_2$, is precipitated. When heated in a current of carbon dioxide, stannous hydroxide forms black stannous oxide, SnO ; which, when heated in air, burns to stannic oxide, SnO_2 . Stannous hydroxide is insoluble in ammonia, but it readily dissolves in alkaline solutions, forming stannites, e.g. potassium stannite: $\text{Sn}(\text{OH})_2 + 2\text{KOH} \rightleftharpoons \text{Sn}(\text{OK})_2 + 2\text{H}_2\text{O}$. Hence the stannites are to be regarded as salts of stannous acid, H_2SnO_2 , which may be simply stannous hydroxide. Basic stannites of the type $\text{HOSn}(\text{OK})$ have been isolated. By the slow evaporation of alkaline solutions of the stannites, crystals of SnO can be obtained. If much alkali be present, or if the solution be boiled, metallic tin separates and alkali stannates, K_2SnO_3 , are formed: $2\text{KHSnO}_2 = \text{K}_2\text{SnO}_3 + \text{Sn} + \text{H}_2\text{O}$. Stannous hydroxide is converted into stannous salts when treated with acids, and consequently, stannous hydroxide exhibits both acidic and basic properties.

§ 4. Stannic Oxide and Hydroxide.

If an acidified solution of stannous chloride, SnCl_2 , be exposed to the air, oxygen is absorbed, and stannic hydroxide is precipitated: $2\text{SnCl}_2 + \text{O}_2 + 6\text{H}_2\text{O} \rightarrow 2\text{Sn}(\text{OH})_4 + 4\text{HCl}$. Possibly stannic chloride is first formed: $2\text{SnCl}_2 + \text{O}_2 + 4\text{HCl} \rightarrow 2\text{SnCl}_4 + 2\text{H}_2\text{O}$; and this latter is hydrolyzed by the water: $\text{SnCl}_4 + 4\text{H}_2\text{O} \rightleftharpoons \text{Sn}(\text{OH})_4 + 4\text{HCl}$. The stannic hydroxide is precipitated by ammonia, alkaline hydroxides, carbonates, ammonium nitrate, sodium sulphate, etc., from solutions of stannic salts: $\text{SnCl}_4 + 4\text{KOK} \rightleftharpoons 4\text{KCl} + \text{Sn}(\text{OH})_4$. Stannic chloride is hydrolyzed by water with the production of the same hydroxide. If the precipitate be dried in air, it has the empirical composition H_4SnO_4 , or $\text{Sn}(\text{OH})_4$; and if dried over concentrated sulphuric acid, the empirical composition H_2SnO_3 , or $\text{SnO}(\text{OH})_2$. Consequently H_4SnO_4 , that is, $\text{Sn}(\text{OH})_4$, is to be regarded as ortho-stannic acid, and H_2SnO_3 , or $\text{O}=\text{Sn}=(\text{OH})_2$, as meta-stannic acid. Stannates corresponding with $\text{K}_2\text{SnO}_3 \cdot 3\text{H}_2\text{O}$, analogous with the carbonates, can be made by heating solutions of, say, stannic chloride with an alkaline carbonate; $\text{SnO}_2 \cdot \text{H}_2\text{O}$ is precipitated from aqueous solutions of the stannates by carbon dioxide.

When metallic tin is treated with hot nitric acid (specific gravity 1.3) stannic hydroxide with the empirical formula $\text{Sn}(\text{OH})_4$ is formed, and this when dried in air has the empirical formula H_4SnO_4 , and if dried over sulphuric acid it has the empirical formula $\text{SnO}(\text{OH})_2$. The stannic acids formed by these two different processes differ essentially in their behaviour towards many reagents. For convenience, the stannic acid formed by the action of alkalis on solutions of stannic chloride is called α -stannic acid,

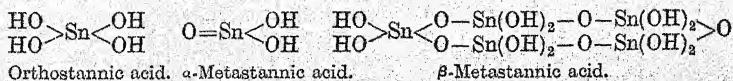
and the acid produced by the action of nitric acid on the metal is called β -stannic acid or "metastannic acid." Some of the differences between the two varieties are as follows :

TABLE LXI.—PROPERTIES OF THE METASTANNIC ACIDS.

α -Metastannic acid.	β -Metastannic acid.
Salts dissolve in water easily and are not decomposed. More basic than the β -acid.	Salts dissolve in water with difficulty and form insoluble basic salts and free acid. Less basic than the α -acid.
When moist, dissolves readily in nitric acid.	Insoluble in nitric acid.
Soluble in dilute sulphuric acid and the solution does not gelatinize when boiled.	Insoluble in sulphuric acid even if concentrated.
Easily soluble in hydrochloric and the solution remains clear when boiled.	Unites with hydrochloric acid forming a substance insoluble in acid but soluble in water. The aqueous solution gelatinizes when boiled.

Both varieties dissolve in caustic alkalis and in alkaline carbonates, and when reprecipitated by the addition of acids, the precipitates retain the properties they had before dissolving in the alkalis. J. J. Berzelius observed this peculiarity of the stannic acids in 1811-1823. Berzelius found that both oxides had the same ultimate composition, and the fact was regarded at the time as an "interesting example," showing that "identity of composition does not correspond with identity of properties." An explanation of the difference between the two acids is not known with certainty. It is generally supposed that the β -acid— $(\text{H}_2\text{SnO}_3)_5$ or $\text{Sn}_5\text{O}_5(\text{OH})_{10}$ —is a polymerized form of the α -acid— H_2SnO_3 . If the β -acid be treated with concentrated hydrochloric acid for a short time, a compound $\text{Sn}_5\text{O}_5\text{Cl}_{10}$ is formed which is insoluble in hydrochloric acid. This substance is soluble in water with the formation of the so-called β -stannic chloride or stannyl chloride $\text{Sn}_5\text{O}_5\text{Cl}_2(\text{OH})_8$. If the aqueous solution be boiled, the β -acid is reprecipitated: $\text{Sn}_5\text{O}_5\text{Cl}_2(\text{OH})_8 + 2\text{H}_2\text{O} = 2\text{HCl} + \text{Sn}_5\text{O}_5(\text{OH})_{10}$; and if the aqueous solution be treated with hydrochloric acid, a precipitate corresponding with $\text{Sn}_5\text{O}_5\text{Cl}_4(\text{OH})_8$ is formed. While tartaric acid prevents the precipitation of stannic hydroxide from stannic chloride solutions by ammonia, the precipitation is not prevented from solutions of β -stannic chloride.

These peculiarities are explained on the assumption that the two isomeric stannic acids have the constitutional formulæ:



When treated with hydrochloric acid, the hydroxyl (OH) groups are replaced by chlorine, and the resulting compound, when treated with water, is more or less completely hydrolyzed. As indicated above, if the β -acid is boiled for a long time with concentrated hydrochloric acid, or concentrated alkaline hydroxide, it is gradually converted into the α -acid; and

conversely, the α -acid is gradually changed into the β -acid at ordinary temperatures, but more rapidly on boiling. Thus: $5\text{SnCl}_4 + 13\text{H}_2\text{O} = 18\text{HCl} + \text{Sn}_5\text{O}_5\text{Cl}_2(\text{OH})_8$. An aqueous solution of β -stannyl chloride, $\text{Sn}_5\text{O}_5\text{Cl}_2(\text{OH})_8$, gives a precipitate of β -stannyl sulphate with sulphuric acid or potassium sulphate. β -stannic acid unites with alkalies, forming β -stannates, e.g. potassium β -stannate, $\text{K}_2\text{Sn}_5\text{O}_{11} \cdot 4\text{H}_2\text{O}$. An excess of alkali precipitates the β -stannates. Hence β -stannic acid exhibits both acidic and basic characteristics. Both stannic acids redden blue litmus, both can be obtained in a colloidal condition by the dialysis of solutions of the corresponding stannates acidified with hydrochloric acid.

An idea is growing in the minds of those who have specially studied the stannic acids that the essential difference between the two varieties is size of grain; that the particles of the α -acid are the smaller and of the β -acid the larger; and that there are not two distinct and well-defined α - and β -stannic acids, since a continuous gradation of forms intermediate between the two varieties can be made by varying the temperature of precipitation. Compare the α - and β -silicic acids, p. 950.

Stannic oxide, SnO_2 .—When the stannic acids are calcined, stannic oxide, SnO_2 , is formed as a white powder which appears yellow when hot. Stannic oxide is also formed as a white powder by the combustion of tin. If stannic chloride be heated in a current of steam, glistening rhombic crystals of stannic oxide are formed isomorphous with brookite, TiO_2 . As previously indicated, stannic oxide occurs in nature in tetragonal crystals of cassiterite which are usually black in colour owing to the presence of impurities. These crystals are isomorphous with rutile, TiO_2 . These crystals also separate on cooling a fused solution of stannic oxide in borax. When stannic oxide is strongly heated with phosphoric acid, crystals of stannic oxide isomorphous with anatase, TiO_2 , are formed. Hence stannic and titanic oxides are isomorphous. Stannic oxide is insoluble in acids and aqueous solutions of the alkalies, but it forms α -alkali stannates when fused with sodium or potassium carbonate or hydroxide. Silica behaves in a similar manner. Stannic oxide melts at about 1625° .

§ 5. Lead.

History.—Lead was known to the ancient Egyptians. It is mentioned several times in the Old Testament. It appears to have been confused with tin (*q.v.*), and Pliny seems to have distinguished between *plumbum nigrum* (black lead) and *plumbum album* or *plumbum candidum*. The ancient Romans used lead for making water-pipes, and some lead compounds were used as cosmetics, and as paint. The alchemists connected lead with the slow-moving planet Saturn, and accordingly represented lead by a scythe ♄ , the symbol for Saturn, because, as some writers say, they thought it to be the oldest of metals and therefore comparable with Saturn, the father of the gods; others say that it has reference to its power of dissolving other metals which recalled the practice of Saturn in devouring his own children. These notions are products of an unbridled imagination.

Occurrence.—Small quantities of metallic lead are occasionally found in nature. In combination with sulphur, lead occurs as sulphide, *galena*,

PbS. This is the most abundant ore of lead. Commercial lead is obtained almost exclusively from galena. Lead carbonate, *cerussite*, PbCO_3 , is not uncommon. Lead sulphate, PbSO_4 , occurs as *anglesite*; lead chromate, *crocosite*, PbCrO_4 ; lead molybdate, *wulfenite*, PbMoO_4 ; lead phosphate, *pyromorphite*, $\text{PbCl}_2 \cdot 3\text{Pb}_3(\text{PO}_4)_2$; lead chloride, *matlockite*, PbCl_2 . Lead ores come from England, United States, Germany, Mexico, Spain, New South Wales, South America, etc.

Extraction.—G. Agricola in his *De re Metallica* (1556) gave a description of the methods of smelting lead ores in vogue among the Saxons, Poles, and Westphalians in the 16th century. Alternate layers of fuel (wood) and ore were piled on a "hearth of bricks—four feet high—with sloping sides," so as to form a kind of basin. When the fuel burnt the ore was reduced, and the resulting metal collected in the basin of the hearth. The lead basin was either tipped like a crucible, or the lead was allowed to flow over the edge at one corner. Open-hearth smelting persists in many places even to-day, although it has yielded to reverberatory or blast furnaces in other places. There are several minor improvements in the modern open-hearth furnace. The lead pot is heated from below, the natural draft has given way to the fan blower, the fumes are collected in chambers to protect the workmen from what has been styled "the crippling fume ghost." When the galena is roasted in a current of air at a low temperature in a reverberatory furnace—say 600° – 900° —so that one part of the lead sulphide, PbS, is oxidized to lead monoxide, PbO, and another part is oxidized to lead sulphate: $2\text{PbS} + 3\text{O}_2 = 2\text{PbO} + 2\text{SO}_2 + 199.7 \text{ Cals.}$; $\text{PbS} + 2\text{O}_2 = \text{PbSO}_4 + 199.5 \text{ Cals.}$ If the mixture of lead sulphide, lead sulphate, and lead monoxide so obtained be heated to a higher temperature with the air "shut off," a series of endothermal reactions then set in. The sulphide reacts with the sulphate and the lead monoxide, forming metallic lead: $\text{PbSO}_4 + \text{PbS} = 2\text{Pb} + 2\text{SO}_2 - 97.4 \text{ Cals.}$; and $2\text{PbO} + \text{PbS} = 3\text{Pb} + \text{SO}_2 - 52.5 \text{ Cals.}$; and towards the end of the operation, when the lead sulphide begins to fuse: $\text{PbS} + 3\text{PbSO}_4 = 4\text{PbO} + 4\text{SO}_2 - 187.1 \text{ Cals.}$ If carbonaceous fuel is present, the lead oxide and lead sulphide are reduced to metallic lead. In dealing with ores low in lead, e.g. some Spanish ores, the ore is sometimes reduced by heating with iron, or a mixture of iron ore and coke. Metallic lead and iron sulphide are formed: $\text{PbS} + \text{Fe} = \text{Pb} + \text{FeS} + 3.8 \text{ Cals.}$

The lead made by these processes usually contains antimony, tin, copper, etc. These impurities make the lead hard and brittle. A large proportion can be removed by heating the metal in a shallow flat-bottomed reverberatory furnace whereby most of the impurities are oxidized before the lead, and rise to the surface as a scum. This operation is known as "softening lead." The silver is usually extracted by Parkes' process.

Properties of lead.—Lead is a bluish-grey metal with a bright metallic lustre when freshly cut, but the lustre soon disappears in ordinary air. Perfectly dry air, and air-free water, have no action on the metal, but if moist air be present, or if the metal be immersed in aerated water, lead is soon covered with a film, probably an oxide, and this is ultimately converted into a basic carbonate. Lead is soft enough to be cut with a knife and scratched with the finger nail. It leaves a grey streak when drawn across paper. Small traces of impurity—antimony, arsenic, copper, zinc—make the lead much harder. Lead is not tough enough to be hammered

into foil or drawn into wire; but it can be pressed into pipes, or rolled into thin sheets or foil. Lead filings under a pressure of about 13 tons per square inch form a solid block; and the metal seems to liquefy under a pressure of about 33 tons per square inch. The specific gravity of lead varies from 11.25 to 11.4 according as the metal is cast or rolled. Lead melts at 326° , and boils at about 1525° . When cooled slowly, the molten metal forms a mass of octahedral crystals (cubic system). The crystalline nature of the metal is shown by the electrolysis of a lead salt. Lead is also deposited as an "arborescent" mass of crystals—called a "lead tree"—when a strip of iron or zinc is suspended in a solution of a lead salt.

Lead is fairly rapidly dissolved by nitric acid, but is little affected by dilute hydrochloric or dilute sulphuric acid in the cold, because a crust of insoluble lead chloride or sulphate is formed on the surface, and this protects the metal from further action. Powdered lead is quickly dissolved by boiling concentrated hydrochloric or sulphuric acid. Organic acids—acetic acid (vinegar)—also act as solvents for metallic lead. Hence vessels plated with tin containing lead, if used for cooking purposes, may contaminate the food with poisonous lead compounds. Water containing sulphates and carbonates in solution forms a coating on the surface of lead which prevents further action. Lead is attacked by water holding air, nitrates, ammonium salts, and carbon dioxide in solution. In the latter case, a soluble acid carbonate may be formed. All soluble lead salts are poisonous, and if the water supply of a town be pure enough to attack lead it may produce lead poisoning. To avoid risk, the water is sometimes filtered through limestone or chalk since carbonates in the water do not corrode the pipes very much. The water then takes up enough carbonates to form a film on the interior of the lead pipes which protects the lead from further action.

Atomic weight of lead.—The combining weight of lead has been determined by the synthesis of lead nitrate and lead sulphate from metallic lead; by the analysis of lead chloride, etc. If oxygen be 16, the best determinations of the combining weight of lead vary between 206.8 and 207.6. The best representative value is supposed to be 207.1. This agrees with the result by Dulong and Petit's method of approximation since the specific heat of lead is 0.0309, and $6.4 \div 0.0309 = 207.1$. The vapour density of the volatile lead compounds also corresponds with the atomic weight 207.1. The vapour density of metallic lead between 1870° and 2000° shows that the molecule is monatomic.

Uses.—Lead is largely used in the arts on account of the ease with which it can be worked, cut, bent, soldered, and on account of its power of resisting attack by water and many acids. It is used in the manufacture of pipes for conveying water; for the manufacture of sheaths for electric wires, sheets for sinks, cisterns, and roofs, lead chambers for sulphuric acid works, evaporation pans in chemical works, etc. It is used in making bullets, shot, accumulator plates, etc. Type metal, solder, pewter, and fusible alloys contain much lead. These alloys have been previously discussed. In soldering, the metal surfaces must be clean and free from oxides, grease, etc. Zinc chloride, "killed spirits," resin, or a little borax is employed either to dissolve the oxide if formed, or to protect the surface from oxidation during the soldering.

§ 6. Lead Monoxide and Hydroxide.

The dull grey iridescent coating which is formed on the surface of lead melted at a low temperature is supposed to be lead suboxide, Pb_2O .¹ If this scum be continuously removed, and again heated to a low temperature in air so as to avoid fusion, a yellow powder of lead monoxide, PbO , sometimes called *massicot* is formed. If this oxide be fused it forms, on cooling, a buff-coloured crystalline mass of lead monoxide, or litharge, PbO . If slowly cooled the oxide has a reddish tint, and if rapidly cooled, a yellow tint. It is thought that there are two modifications of lead monoxide—yellow and red, the latter being the stable form at ordinary temperatures. Litharge is commonly made by the cupellation of lead. Large quantities of lead monoxide are used in preparing oils, and varnishes, and in making flint glass.

Lead monoxide fuses at 877° , and it is volatile at a red heat. Its specific gravity is 9.36. Lead monoxide is slightly soluble in water, probably forming lead hydroxide, $\text{Pb}(\text{OH})_2$; sufficient lead hydroxide is dissolved to give the water an alkaline reaction. Lead hydroxide is best made by the addition of alkali to a lead salt. The composition of the precipitate depends upon the conditions of precipitation. The hydroxides, $2\text{PbO} \cdot \text{H}_2\text{O}$ and $\text{PbO} \cdot \text{H}_2\text{O}$, have been reported. The precipitate is soluble in an excess of alkali hydroxide, forming plumbites. For instance, potassium plumbite, $\text{Pb}(\text{OK})_2$, is formed when potassium hydroxide is used. The precipitate is insoluble in ammonia. When heated to 145° , lead hydroxide passes into the oxide. Both oxide and hydroxide dissolve in acids, forming lead salts. The formation of plumbites illustrates the acidic nature of lead monoxide; and the formation of lead salts, and the alkalinity of aqueous solutions of lead hydroxide, show the basic properties of this oxide. Hence lead monoxide is an amphoteric oxide.

§ 7. Lead Sesquioxide and Red Lead.

Lead sesquioxide, Pb_2O_3 .—This oxide is an orange-yellow powder formed when lead hydroxide, in alkaline solution, is treated with an oxidizing agent—hypochlorites, chlorine, bromine, hydrogen peroxide, potassium persulphate— $2\text{PbO} + \text{NaOCl} = \text{NaCl} + \text{Pb}_2\text{O}_3$. When heated lead sesquioxide forms lead monoxide and oxygen. Acids, say nitric acid, decompose it into lead monoxide and lead dioxide—the former dissolves in the acid, forming lead nitrate. With concentrated hydrochloric acid, lead sesquioxide behaves like a peroxide furnishing chlorine gas: $\text{Pb}_2\text{O}_3 + 6\text{HCl} = 3\text{H}_2\text{O} + 2\text{PbCl}_2 + \text{Cl}_2$.

Lead tetroxide, minium, red lead, Pb_3O_4 .—Red lead is formed by

¹ The same oxide is made by heating lead oxalate to about 300° in a glass retort: $2\text{PbC}_2\text{O}_4 = \text{CO} + 3\text{CO}_2 + \text{Pb}_2\text{O}$. A current of carbon dioxide is directed through the retort during the operation, or the lead suboxide will be partly reduced by the carbon monoxide formed during the reaction. This oxide burns to PbO when heated in air, and if heated to about 350° , the suboxide decomposes into a greenish-grey powder which is a mixture of lead monoxide and metallic lead. Lead suboxide is insoluble in and unchanged by water; it is decomposed by alkalis and acids; it is not altered by exposure to dry air; its specific gravity is 8.348; and the thermal value of its reaction with acetic acid is not in favour of the assumption that it is a mixture of lead and lead monoxide: $\text{Pb} + \text{PbO}$.

heating lead monoxide or white lead in air between 470° and 480° . It is a scarlet powder. The hot powder acquires a deeper and deeper tint as time goes on, then it appears to become violet, and finally black. On cooling the original red powder appears. The tint of red lead is determined largely by the mode of preparation; its composition also varies with the temperature and time of heating. Although there is little doubt that a compound of the empirical formula Pb_3O_4 does exist, most red leads of commerce appear to contain Pb_3O_4 with variable amounts of lead monoxide. Conversely, litharge generally contains a little red lead. Many varieties of red lead approximate in composition to Pb_4O_5 , and these red leads no doubt contain $Pb_3O_4 + nPbO$. As with lead sesquioxide, dilute acids decompose Pb_3O_4 into PbO and PbO_2 , the former dissolves in the acid, forming a salt of lead. When heated, red lead dissociates into lead monoxide and oxygen: $2Pb_3O_4 \rightleftharpoons 6PbO + O_2$. The dissociation pressure varies with the temperature, thus:

Temperature	. . . 445°	500°	555°	636°
Pressure	. . . 5	60	183	763 mm.

The partial pressure of oxygen in air is one-fifth of 760 mm., that is 152 mm. This means that red lead will decompose at about 550° when heated in air; and in pure oxygen at atmospheric pressure, red lead can be formed at 600° , but not at 650° . Red lead proper is not a mixture of lead monoxide and lead peroxide because the dissociation pressure of lead peroxide at a given temperature is much less than that of red lead. Red lead is used in the manufacture of flint glass and glazes for pottery. It also is used in the manufacture of paint and of matches. The term *minium* often used for red lead was formerly applied to cinnabar. Cinnabar was and is sometimes adulterated with red lead.

§ 8. Lead Peroxide.

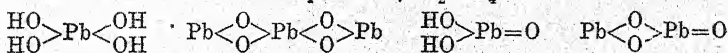
Lead peroxide is prepared by the action of oxidizing agents—hypochlorites, chlorine, bromine, hydrogen peroxide, or persulphates—on alkaline solutions in which lead monoxide is suspended. The same compound is deposited on the anode when a solution of a lead salt is electrolyzed. Lead peroxide is a strong oxidizing agent. Thus, when lead peroxide is gently rubbed with sulphur on a warm surface the mass inflames; with phosphorus, the mixture detonates; when sulphur dioxide is passed over lead dioxide, the two unite to form lead sulphate, $PbSO_4$; and when hydrogen sulphide is allowed to impinge on a few grams of lead peroxide, the gas is oxidized and bursts into flame.

Lead dioxide dissolves in a boiling concentrated aqueous solution of potassium hydroxide, and the solution on cooling deposits crystals of potassium metaplumbate, $K_2PbO_3 \cdot 3H_2O$, analogous with potassium stannate, $K_2SnO_3 \cdot 3H_2O$. If a mixture of calcium carbonate and lead monoxide be heated to 700° , carbon dioxide is given off and oxygen is absorbed, forming calcium metaplumbate, $CaPbO_3$. If calcium plumbate be treated with carbon dioxide at the same temperature, the reaction is reversed, and oxygen is evolved. This process was proposed by G. Kassner for the preparation of oxygen; and more recently, Kassner (1912) proposed a mixture of sodium manganate and metaplumbate for preparing oxygen. This

mixture, with the trade-name *plumboxan*, gives off oxygen when heated in a current of steam; and the plumboxan is restored when heated in a current of air.

The formation of calcium metaplumbate shows that lead dioxide possesses feeble acidic properties. Lead in lead dioxide is probably quadrivalent as in lead tetrachloride and lead tetrafluoride— PbCl_4 and PbF_4 . The existence of quadrivalent lead is established in organic chemistry where a whole series of compounds of the type $\text{Pb}(\text{CH}_3)_4$ are known—this particular one, for instance, is a liquid boiling at 110° , and has a vapour density in harmony with the formula given. Lead dioxide behaves as a feeble base with acetic acid. When red lead is dissolved in hot glacial acetic acid, no oxygen is evolved and pale green needle-like crystals of lead tetracetate separate on cooling the solution. The crystals melt at 175° . The salt is immediately decomposed by water and brown lead peroxide is precipitated; similar evidence for lead tetraphosphate, $\text{Pb}_3(\text{PO}_4)_4$, has been obtained.

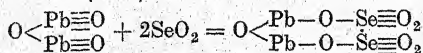
The constitution of the higher oxides of lead.—Lead peroxide is a polyoxide $\text{O}=\text{Pb}=\text{O}$ because it gives oxygen, not hydrogen peroxide, when treated with acids. Thus with concentrated sulphuric acid: $2\text{PbO}_2 + 2\text{H}_2\text{SO}_4 \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O} + \text{O}_2$; and with concentrated hydrochloric acid, it gives chlorine: $\text{PbO}_2 + 4\text{HCl} \rightarrow 2\text{H}_2\text{O} + \text{PbCl}_2 + \text{Cl}_2$. Lead peroxide is best regarded as the anhydride of orthoplumbic acid, $\text{Pb}(\text{OH})_4$, or metaplumbic acid, $\text{PbO}(\text{OH})_2$, just as CO_2 , SiO_2 , SnO_2 , and MnO_2 are the anhydrides of carbonic, silicic, stannic, and manganous acids. When hypochlorites, or other oxidizing agents, act upon an alkaline solution of lead monoxide, a brown precipitate of metaplumbic acid is formed: $\text{Pb}(\text{OH})_2 + 2\text{NaOH} + \text{Cl}_2 \rightarrow \text{H}_2\text{O} + 2\text{NaCl} + \text{PbO}(\text{OH})_2$. If the latter be warmed, it passes into the anhydride PbO_3 . The other two oxides of lead— Pb_2O_3 and Pb_3O_4 —are in all probability salts of plumbic acid; lead sesquioxide appears to be lead metaplumbate, PbPbO_3 ; and red lead to be a lead orthoplumbate, Pb_2PbO_4 .



Orthoplumbic acid, Lead orthoplumbate, Metaplumbic acid, Lead metaplumbate,



The salt, $\text{Pb}_2\text{O}_3(\text{SeO}_2)_2$, or $\text{Pb}_2\text{Se}_2\text{O}_7$, formed by the action of lead sesquioxide on selenious acid, shows that lead sesquioxide can act as a feebly basic oxide. The alternative constitution of lead sesquioxide suggested by L. Marino (1905) is:



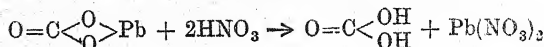
Thallic oxide, Tl_2O_3 , forms a similar compound.

Calcium orthoplumbate, Ca_2PbO_4 , is a crystalline salt analogous with red lead, that is, lead orthoplumbate; and potassium metaplumbate, K_2PbO_3 , is analogous with lead sesquioxide, that is, with lead metaplumbate.

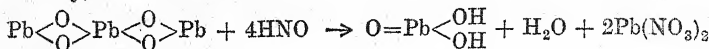
When lead sesquioxide is treated with nitric acid, it forms brown metaplumbic acid, $\text{PbO}(\text{OH})_2$, or its anhydride, PbO_3 :



analogous with the action of nitric acid on lead carbonate :



Similarly, with red lead :

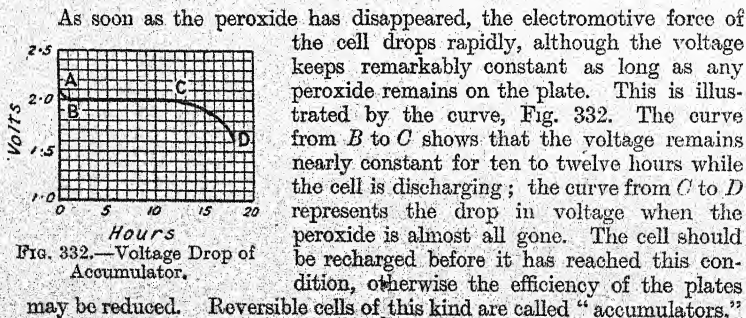
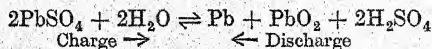


§ 9. Accumulators or Storage Cells

The electric storage battery does not really store electric energy, but it stores energy by the conversion of the electrical into chemical energy, and re-conversion, in the discharge, of the chemical into electrical energy.—C. P. STEINMETZ (1914).

An accumulator is essentially a voltaic cell with plates of lead and lead dioxide with dilute sulphuric acid as the electrolyte: $\text{Pb} \mid \text{H}_2\text{SO}_4 \mid \text{PbO}_2$. If two corrugated lead plates be covered with a paste of litharge, PbO , and dipped in a 20 per cent. solution of sulphuric acid, the litharge on the plates is converted into lead sulphate, PbSO_4 . If an electric current be passed through the cell, hydrogen is evolved at the cathode, and oxygen at the anode during the electrolysis of the sulphuric acid. The hydrogen at the cathode reduces the lead sulphate producing a grey film of "spongy" metallic lead: $\text{PbSO}_4 + \text{H}_2 = \text{H}_2\text{SO}_4 + \text{Pb}$; and the oxygen at the anode transforms lead sulphate into lead persulphate: $2\text{PbSO}_4 + 2\text{H}_2\text{SO}_4 + \text{O}_2 = 2\text{Pb}(\text{SO}_4)_2 + 2\text{H}_2\text{O}$; and the persulphate is at once hydrolyzed into lead peroxide and sulphuric acid: $\text{Pb}(\text{SO}_4)_2 + 2\text{H}_2\text{O} = \text{PbO}_2 + 2\text{H}_2\text{SO}_4$, with the result that a dark brown film of lead peroxide is formed on the lead plate. The current may then be stopped.

If the terminals of the cell be then connected with a suitable resistance, a current of an intensity of nearly two volts can be obtained. The time of discharge—the ampère-hour capacity—assuming the resistance to be constant, depends on the area of the plates. The cell is then said to be "discharged." The brown film of lead peroxide on the anode gradually disappears during the discharge of the cell, and a white film of lead sulphate takes its place. The reaction on the anode during the discharge of the cell is $\text{PbO}_2 + \text{Pb} + 2\text{H}_2\text{SO}_4 = 2\text{PbSO}_4 + 2\text{H}_2\text{O}$. When the cell is being charged, the reaction can be represented by the same equation taken from right to left :



"secondary cells," or "storage cells." An accumulator is thus a cell with metallic lead plates, and the plates specially designed to hold as large an amount of litharge, or lead oxide, as possible. The plates dip in dilute sulphuric acid. When charged, the accumulator acts as if it were a cell $\text{PbO}_2 \mid \text{H}_2\text{SO}_4\text{aq} \mid \text{Pb}$. To summarize the changes at the electrode during charge and discharge :

	Charging.	Discharging.
Cathode . .	$\text{PbSO}_4 \rightarrow \text{Pb}$	$\text{Pb} \rightarrow \text{PbSO}_4$
Anode . .	$\text{PbSO}_4 \rightarrow \text{PbO}_2$	$\text{PbO}_2 \rightarrow \text{PbO} \rightarrow \text{PbSO}_4$

During the discharge, the surface of both plates becomes covered with a film of white lead sulphate, and hydrogen is evolved at the cathode: $\text{Pb} + \text{H}_2\text{SO}_4 = \text{PbSO}_4 + \text{H}_2$. The hydrogen is partly absorbed by the lead which has not been converted into sulphate. The presence of the absorbed gas raises the electromotive force of the cell during discharge until it is all consumed. This occupies but a few minutes ; during this time, the voltage is represented by the curve *AB*.

Several other interpretations of the action in an accumulator during charge and discharge have been proposed. The hypothesis indicated above can be easily translated into the language of the ionic hypothesis. In the charged cell, we have a liquid containing H^+ and SO_4^- ions, with lead and lead peroxide electrodes. When the cell is discharging, the H^+ ions travel towards the PbO_2 plate and reduce the lead peroxide to lead monoxide: $2\text{H}^+ + \text{PbO}_2 = \text{H}_2\text{O} + \text{PbO}$; and this plate accordingly receives a positive charge, owing to the de-electrification of the H^+ ions. The SO_4^- ions simultaneously travel to the lead plates converting the lead at the surface of the plate into lead sulphate: $\text{SO}_4^- + \text{Pb} \rightarrow \text{PbSO}_4$, and the plates are at the same time charged negatively. In consequence, a positive current of electricity travels from the "peroxide" plate to the "lead" plate outside the cell, and from the "lead" to the "peroxide" plate in the liquid. A secondary reaction between the sulphuric acid and the lead monoxide of the "peroxide" plate leads to the formation of lead sulphate, and consequently the consumption of the sulphuric acid in the liquid in the cell is relatively large during discharge. When the cell is to be re-charged, the "peroxide" plate is connected with the positive pole of the dynamo, and the "lead" plate with the negative pole of the dynamo. The H^+ ions of the cell travel to the negatively charged plate, an equivalent number of SO_4^- ions pass into the solution and metallic lead remains on the plate. The effect is to reduce the lead sulphate back to metallic lead. The SO_4^- ions travel to the positively charged "peroxide" plate and form lead persulphate: $2\text{SO}_4^- + 2\text{PbSO}_4 = 2\text{Pb}(\text{SO}_4)_2$. The lead persulphate is immediately hydrolyzed to lead peroxide as indicated above. Sulphuric acid is thus regenerated, and the cell is brought back to its original condition.

Numerous attempts to find a substitute for the heavy lead plates of the Planté or Faure storage cell have met with but little success. The most promising is the so-called iron-nickel cell: $\text{Fe} \mid \text{KOH aq} \mid \text{NiO}_2$. The chemistry of this cell turns on the reaction: $\text{NiO}_2 + \text{Fe} \rightleftharpoons \text{NiO} + \text{FeO}$.

§ 10. The Relationships of the Carbon-Silicon-Tin Family.

The elements carbon, silicon, germanium, tin, and lead have a close family relationship. The physical properties, it will be seen, where known, vary with the atomic weight. Thus :

TABLE LXII.—PROPERTIES OF THE CARBON-TIN FAMILY.

	Carbon.	Silicon.	Germanium.	Tin.	Lead.
Atomic weight . .	12	28.3	72.5	119	207.1
Specific gravity . .	2.3 to 3.5	2.35	5.47	5.8 to 7.3	11.4
Atomic volume . .	4.15	12.04	13.26	18.25	18.18
Melting point . .	—	1350°	958°	231°	326°
Boiling point . .	—	about 3000°	—	2200°	1500°

The gradual change in the physical properties with rise in atomic weight is very manifest. The chemical properties have many interesting resemblances. They are all bi- and quadri-valent, and there is a marked transition from acidic to basic qualities with rise of atomic weight. Carbon and silicon both form hydrogen compounds and exhibit some remarkable analogies. The relationships of the members of this series are best worked out by comparing the properties of the halides, the monoxides and peroxides, the occurrence, the allotropic forms of the elements, etc. The monoxides of silicon and of the remainder of these elements are solid, carbon monoxide is gaseous. There are a large number of stannous and plumbous salts corresponding with the monoxides of tin and lead. The germanous chloride, GeCl_2 , is a liquid hydrolyzed by water; stannous chloride, SnCl_2 , is more stable, but it too is hydrolyzed by water; and lead chloride is stable in water. The -ic salts of germanium are more stable than the -ous salts, whereas with lead the converse is the case. The dioxides are obtained by strongly heating the elements in air, but in the case of lead, the conversion is incomplete since PbO_2 is not the stable form of lead oxide. Silica, SiO_2 , dissolves in hydrofluoric acid; germanic oxide, GeO_2 , in acids; and both tin and lead dioxides dissolve in acids to form salts. The dioxides are distinctly acidic, uniting with alkalis to form salts of the type MRO_3 . The elements all form sulphides of the types MS and MS_2 , although the existence of PbS_2 is not well established. The mono sulphides GeS , SnS , and PbS , react with alkalis to form thiosalts of the type R_2MS_3 . These elements all form chlorides of the type MCl_4 . Like bismuth in the phosphorus family, lithium in the alkali family, mercury in the zinc family, lead seems to link this family with some other family group.

§ 11. Titanium, Zirconium, Cerium, and Thorium.

	Titanium, Ti.	Zirconium, Zr.	Cerium, Ce.	Thorium, Th.
Atomic weight . .	48.1	90.6	140.25	232.4

Titanium.—Titanium was discovered in 1791 by W. Gregor while investigating the magnetic sand (menachanite) found in Menachan (Cornwall). He called this element "menachin." Three years later, M. H.

Klaproth found what he supposed to be a new earth in rutile. He called the metal derived from the earth "titanium," and, in 1797, he showed that titanium was identical with the menachin of W. Gregor. The term "titanium" was derived from "Titans," the fabled giants of ancient mythology. J. J. Berzelius first isolated the metal—more or less impure—in 1825.

Titanium does not occur in nature free; but combined, small quantities are exceedingly common. Most sands, clays, granitic rocks contain a small proportion, say, about 0.5 per cent. Titanium has been detected in many mineral waters, in many plants, the bones of animals, in the atmosphere of the sun (by the spectroscope), etc. The chief minerals are *rutile*, TiO_2 (tetragonal, sp. gr. 4.3), and the trimorphic *brookite*, TiO_2 (orthorhombic, sp. gr. 4.0), and *anatase*, TiO_2 (tetragonal, sp. gr. 3.9); titanium also occurs in minerals *ilmenite* (titaniferous iron ore), FeTiO_3 ; *sphene* or *titanite*, or calcium titanium silicate, CaTiSiO_5 , that is $\text{CaO} \cdot \text{TiO}_2 \cdot \text{SiO}_2$.

We have just seen that titanic oxide, TiO_2 , is trimorphous, and isotrimorphous with stannic oxide, p. 972. Anatase crystals appear to separate when a solution containing titanic oxide is heated to a temperature of about 860° , at about 1000° , crystals of brookite separate, and at higher temperatures, rutile. The isomorphism of tin and titanium oxides is illustrated by the following scheme:

	Tetragonal.	Rhombic.	Tetragonal.
TiO_2	Rutile	Brookite	Anatase
SnO_2	Cassiterite	Artificial	Artificial
SiO_2	—	Tridymite	—

Titanic oxide, or titanium dioxide, TiO_2 , is both acidic and basic. It forms titanates when fused with alkalis, e.g. potassium titanate, K_2TiO_3 , when fused with potassium hydroxide. It also forms salts, TiCl_4 , $\text{Ti}(\text{SO}_4)_2$, etc., where titanium is a tetrad. Titanic chloride, TiCl_4 , is made by passing chlorine over a mixture of titanic oxide and carbon. Titanic hydroxide, $\text{Ti}(\text{OH})_4$, is precipitated from its salts by the addition of alkaline hydroxide or ammonia. When dried it appears to form metatitanic acid, H_2TiO_3 , by the loss of a molecule of water. There is a doubt about the actual formation of a definite acid, but in any case, there is a well-defined series of metatitanates— K_2TiO_3 , FeTiO_3 , etc. Several basic salts have been reported—e.g. $\text{TiO}(\text{SO}_4)$, etc. Metallic zinc or tin reduces acidified solutions of titanic oxide to violet-coloured titanous chloride, TiCl_3 , where titanium is a triad. The corresponding oxide is titanium sesquioxide, Ti_2O_3 . Titanous chloride is a powerful reducing agent—ferric salts are reduced by it to ferrous salts; cupric to cuprous salts, etc. $\text{FeCl}_3 + \text{TiCl}_3 = \text{TiCl}_4 + \text{FeCl}_2$. The reaction with ferric salts is quantitative and takes place in the cold; a standard solution of titanous chloride furnishes a ready means of ascertaining the amount of ferric oxide in a solution without the need for a preliminary reduction of the ferric salt as in the case where a permanganate or dichromate solution is used for the titration. The standard solution of titanous chloride must be protected from oxidation after it has been standardized. The titanous chloride solution is made by the electrolytic reduction of titanic chloride, or by reducing the latter with zinc or tin and acid. The titanic salts are not reduced by hydrogen sulphide

or sulphurous acid, but, like the ferric salts, acidified solutions are reduced to titanous salts by zinc, tin, etc. Hydrogen peroxide produces an orange-yellow coloration in the presence of titanic salts, and this test is used for the colorimetric determination of titanium in rocks, etc. The intensity of the colour is proportional to the amount of titanium present owing, it is supposed, to the formation of a titanium peroxide, TiO_3 . Vanadium salts in a similar manner produce a brick-red coloration. The yellow colour with titanium is supposed to depend upon the formation of titanium trioxide, TiO_3 , where titanium appears to be sexivalent. A series of bivalent titanium salts, blue in colour, are known. The corresponding oxide is titanium monoxide, TiO . The metal is made by reducing the dioxide mixed with carbon in an electric arc furnace, and also by Goldschmidt's process. A sample of the metal of 99.9 per cent. purity melts at 1795° , and burns when heated in oxygen, forming titanium dioxide; and at 800° , in nitrogen, it forms titanium nitride, Ti_3N_4 . The metal decomposes boiling water, and dissolves in acids forming titanic salts. Titanium is closely related with silicon.

Ferro-titanium (an alloy of iron with, usually, 10 to 15 per cent. of titanium) is used in the purification of Bessemer's and the open-hearth steel. Enough titanium is added to give a steel containing 0.05 to 0.2 per cent. of titanium. It is said that the beneficial effect of titanium is not that of an alloying element proper, but that "the titanium acts simply as a purifier, for it is a powerful deoxidizer of molten steel and it removes part if not all the nitrides as well." It prevents brittleness, and allows the steel to carry a greater proportion of carbon, thus improving the durability and toughness of the metal. Titanium carbide has been used as the negative pole for arc-lighting, the positive pole being a rod of copper, or carbon. One of the electrodes in the so-called "magnetite" arc lamp is a mixture of magnetite and chromite with 15 to 20 per cent. of rutile.

Zirconium.—While M. H. Klaproth was analyzing the mineral zircon, in 1788, he found one variety contained :

Silica.	Nickel and iron oxides.	New earth.	
31.5	0.5	68.0	per cent.

Unlike aluminium, the new earth did not dissolve in alkalis. The fact that the new earth was separated from zircon led to the term "zirconia," and "zirconium" for the corresponding metal, which was isolated by J. J. Berzelius in 1824. Several chemists have reported that zirconia is a mixture of two distinct earths: one zirconia proper, and the other a new earth styled by L. F. Svanberg (1845), "noria," and by H. C. Sorby (1869), "jargonite." These pseudo-discoveries have been traced to the contamination of zirconia with known elements—*e.g.* uranium, etc.

Zircon is said to be a silicate of zirconium, ZrSiO_4 , because the ratio of the percentages of silica and zirconia obtained by analysis of the mineral approximate (1:1). The isomorphism of the mineral with tetragonal silica has led some to regard zircon as an isomorphous mixture of the two oxides. Zirconium forms two series of oxides, ZrO_2 and Zr_2O_5 . When zircon is fused with sodium carbonate, and the fused mass is treated with water, sodium zirconate, Na_2ZrO_4 , and sodium silicate, Na_2SiO_4 , pass into solution; the sodium zirconate is immediately hydrolyzed into insoluble zirconium hydroxide, Zr(OH)_4 . This latter, on ignition, furnishes zirconium dioxide

or zirconia, ZrO_2 , which has both acidic and basic properties. The metal is made by reducing the dioxide with carbon in the electric furnace. Zirconia is used in place of lime in Drummond's lamp—zircon lamp. Zirconia mixed with magnesia, thoria, etc., is used as the glower filament in Nernst's lamp. The filament is a non-conductor of electricity at ordinary temperatures, but when heated, it conducts and radiates a greater amount of light for a given consumption of electrical energy than the carbon filament lamp in common use. Zirconia is also used in the manufacture of incandescent mantles, and in the preparation of very refractory crucibles, etc.

Cerium.—In 1803 M. H. Klaproth discovered a new light-brown earth while analyzing a mineral from Ridderhyttan (Sweden). Klaproth called the earth "ochroite," from the Greek $\chi\chi\chi\chi$ (ochros), brownish-yellow. At the same time, and independently of Klaproth, J. J. Berzelius and W. Hisinger made the same discovery, and named the earth "ceria" in honour of the discovery of the planet Ceres by G. Piazzi, in 1801. The term "ceria" has been retained. The element was isolated by Mosander in 1826. Ceria is one of the so-called rare earths (*q.v.*); it occurs more particularly in cerite, euxenite, and monazite. Cerium forms two oxides and two well-defined series of salts in which cerium is respectively ter- and quadri-valent. Both oxides are basic. The former furnishes white cerous salts— $Ce(NO_3)_3$, etc.; the latter, orange-red ceric salts— $Ce(NO_3)_4$, etc. An alloy of cerium and iron—called *Auer metal*—produces bright sparks when drawn across a rough surface, and it is hence used in the manufacture of gas lighters. Cerium metal readily oxidizes on exposure to air. Its specific gravity is 7.04, and it melts at 623° .

Thorium.—In 1818 J. J. Berzelius believed that he had discovered a new earth in a mineral from Fahlun (Sweden), and he gave it the name "thoria"—from Thor, son of the Scandinavian god Thor. There was some doubt at the time about the novelty of this earth. In 1828 Esmark discovered a mineral near Brevig (Norway) from which Berzelius isolated an earth very similar to that previously called "thoria." Esmark's mineral was subsequently called "thorite." The earth thoria was afterwards detected in many other minerals—thorianite (Ceylon), monazite, orangeite, orthite, euxenite, etc. In 1862 J. F. Bahr thought that he had discovered a new element in a mineral resembling orthite. This element was called "wasium." Bahr's wasium was afterwards identified with thorium. Thorium is closely analogous in properties with zirconium and silicon. Thoria is largely used in the manufacture of incandescent mantles (p. 924). It melts and begins to volatilize at about 2000° . The main supply is monazite sand which contains 1 to 7 per cent. of thorium oxide, as well as cerium oxide, etc. The powdered mineral is digested with sulphuric acid. The solution is evaporated to dryness and extracted with water. The solution is treated with oxalic acid when the rare earths and thoria are precipitated as oxalates. Several methods are available for separating thorium, thus, (1) the oxalates can be digested in ammonium oxalate in which thorium oxalate is soluble; and (2) the oxalates can be converted to sulphates and dissolved in ice-cold water. On warming, thorium sulphate crystallizes out.

Family relationship of the group.—There are some marked dissimilarities in the elements of this group, both in physical and chemical properties. This is more particularly the case with cerium, which, in some

respects, can be said to occupy an anomalous position. To summarize the physical properties of the related elements titanium, zirconium, cerium, and thorium:

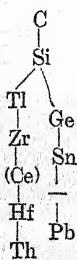
TABLE LXIII.—PROPERTIES OF THE TITANIUM-THORIUM FAMILY.

	Titanium.	Zirconium.	Cerium.	Thorium.
Atomic weight . . .	48.1	90.6	140.25	232.42
Specific gravity . . .	4.87	6.4	7.04	11.00
Atomic volume . . .	9.88	14.1	19.92	21.13
Melting point . . .	1795°	2350°	830°	2000°

The relationship of this group of elements to the carbon-lead family is sometimes represented by means of a diagram like that shown in the margin. The differences between cerium and thorium, and between tin and lead are supposed to indicate the existence of elements not yet discovered, and which are represented in the diagram by hyphens. Titanium is the only element in the group possessing salts which bear a marked resemblance to those of stannous tin. These elements all form dioxides when heated in air or oxygen; and they are all precipitated as hydroxides by the addition of ammonia to solutions of their salts. They all exhibit basic properties uniting with acids to form salts; and when fused with alkalis, titania and zirconia show basic properties forming, say, sodium titanate, Na_2TiO_3 , and sodium zirconate, Na_2ZrO_3 ; thus recalling sodium carbonate, Na_2CO_3 , and sodium silicate, Na_2SiO_3 .

Ceria and thorina do not form similar compounds. The fact that when the hydrated dioxides are precipitated by potash-lye, the whole of the alkali cannot be removed by washing may be a case of adsorption, or it may indicate the possible existence of cerates and thorates. The tetrafluorides of all four elements are formed by dissolving the dioxides in hydrofluoric acid, the fluorides of titanium, zirconium, and thorium unite with the alkali fluorides forming the fluotitanates, fluozirconates, and fluothorates. Cerium forms a double fluoride which is not the same type. Titanium and cerium form an -ous and an -ic series of salts. Zirconium and thorium form -ic salts. The salts of the whole family are very prone to double-salt formation. While carbon tetrafluoride is an inert substance resembling the paraffins, the fluorides of the other elements—silicon, etc.—have a marked residual affinity, which is shown by their forming complex salts like M_2SiF_6 , K_2ZrF_6 , $(\text{NH}_4)_2\text{SnCl}_6$, $(\text{NH}_4)_2\text{PbCl}_6$, $(\text{NH}_4)_2\text{Ce}(\text{NO}_3)_6$, etc. The element hafnium, Hf, recently discovered by D. Coster and G. von Hevesy (1923) to be associated with most zirconium minerals, has an atomic weight about 174.1, and is closely related to zirconium.

The French chemists claim that the name of hafnium should be celtium, because G. Urbain announced the discovery of a new element "celtium" in some rare earth residues. In 1911 celtium proved to be impure lutecium, but in the spectrums of some residues afterwards (1922) obtained there were lines which corresponded with a new element. Our knowledge of the *neoceltium* is very meagre compared even with the little known about hafnium, though both are probably the same element.



§ 12. The Rare Earths.

The rare earths form a group to themselves; chemically, they are so much alike that it taxes the utmost skill of the chemist to effect even a partial separation, and their history is so obscure that we do not yet know the number of them.—W. CROOKES.

In 1794 J. Gadolin discovered a new earth in the mineral gadolinite, which had been named in honour of himself; and in 1797 A. G. Ekeberg named the new earth *yttria*, after Ytterby, the place in Sweden where the mineral was found. C. G. Mosander (1843) noticed that what was then called *yttria* is really a complex earth containing *yttria* proper, and two other earths, which were oxides of two new elements, *erbium* and *terbium*. In 1880 P. T. Cleve found *erbium* earth not only contained *erbium*, but also a new element *thulium*—a name derived from the old Latin *Thule* for the uttermost north—and *holmium*. In 1905 G. Urbain also discovered *lutecium* in the *yttria* earths (p. 781). In 1886 Lecoq de Boisbaudran separated *dysprosium*—from the Greek *δυσπρος* (*dyspros*), difficult, in allusion to the trouble involved in its isolation—from the same earth. In the same year, C. Marignac and Lecoq de Boisbaudran showed that *terbium* earth contained a new element *gadolinium* as well as *terbium*. In 1878 C. Marignac found the oxide of a new element in *gadolinite*. This he named *ytterbium*, after Ytterby (Sweden). In 1879 L. F. Nilson, while extracting *ytterbia* from the mineral *euxenite*, also separated an unknown earth which he called *scandium*, in honour of his country—Scandinavia—where the mineral was found. In 1839 C. G. Mosander separated the oxide of a new element from an earth which had hitherto been thought to be *ceria*. This earth he named *lanthanum*, *λανθάνειν* (*lanthanein*), to hide. In 1841 C. G. Mosander separated another earth—*didymia*—from *lanthanum*. The corresponding element was called *didymium*—from the Greek *διδυμος* (*didymos*), a twin, a name suggested by its close relationship and almost invariable occurrence with *lanthanum*.

The rare earths include a curious group of basic oxides of elements which resemble one another so closely that they appear to form a series of compounds with properties which change but slightly from member to member. The term *rare* has really changed its significance when applied to these earths because some of them are not so scarce and rare as was formerly supposed. The rare earths are found in the minerals mentioned under *cerium*, as well as in numerous other rare and scarce minerals. The rare earths are separated by first digesting the mineral with concentrated sulphuric acid. The clear filtrate is treated with hydrogen sulphide to remove copper, bismuth, etc. The solution of the sulphates is treated with oxalic acid. The precipitate containing the oxalates of the rare earths is then separated into three main groups according to the solubility of their double salts with potassium sulphate. Thus:

Insoluble oxalates.

Insoluble double sulphates		Soluble double sulphates			
		Moderately soluble	Very soluble		
Scandium, Sc . .	44.1	Europium, Eu . .	152.0	Yttrium, Y . . .	89.0
Cerium, Ce . .	140.25	Gadolinium, Gd . .	157.3	Dysprosium, Dy . .	162.5
Lanthanum, La . .	139.0	Terbium, Tb . . .	159.2	Holmium, Ho . . .	163.5
Praseodymium, Pr	140.6			Erbium, Er . . .	167.7
Neodymium, Nd	144.3	•		Thulium, Tm . . .	168.5
Samarium, Sa . .	150.4	•		Neoytterbium, Yb .	172.0
		•		Lutecium, Lu . . .	175.0

Fractional precipitation.—It is very difficult to further separate the members of these series one from the other, and there is some uncertainty as to whether some so-called elements are really homogeneous. The elements of the rare earths are in many cases so closely related that they can only be separated into parts which give evidence of chemical individuality by very special and laborious methods. The fractional separation of these elements depends on a reaction in which there is a difference in the behaviour of the elements, even though the difference be slight. Suppose, for instance, two earths differ slightly in basicity. Add a weak solution of ammonia to a very dilute solution of the earths. The dilution must be such that a turbidity appears only after the lapse of a considerable time. The liquid is then filtered even while the precipitation is still incomplete. There is now a slight difference in the basicity of the two portions of the earth so separated, for the portion still in solution is ever so slightly more basic than the portion precipitated by the ammonia. The process is repeated on the two portions and the separated portions are again treated and the operations are so continued that the differences accumulate systematically and finally become perceptible by physical or chemical tests. The reaction most suitable for the fractional separation depends on the nature of the earths to be treated. For example, fractional precipitation of the hydroxides by means of ammonia when the less soluble earths are deposited first; fractional crystallization of salts—bromates, oxalates, formates, etc.

By fractional precipitation, Lecoq de Boisbaudran, in 1879, separated samarium from didymium (discovered by C. G. Mosander in 1841). Samarium was named after the mineral samarskite which, in turn, was named after a Russian, M. Samarsky. E. Demarçay, in 1901, separated europium from samarium; the residual didymia was supposed to be the oxide of a distinct metal, didymium, with a definite atomic weight and which furnished salts whose aqueous solutions gave a characteristic absorption spectrum. In 1885, by the fractional crystallization of a nitric acid solution of didymium ammonia-nitrate, A. von Welsbach separated this salt into two other salts, one bright green and the other amethyst blue. The corresponding oxides were respectively pale green, and pale blue. The metal derived from the pale green oxide was called praseodymium—from the Greek *πράσινος* (*prasinos*), leek green; and the other was called neodymium—from the Greek *νέος* (*neos*), new. Aqueous solutions of the two fractions have distinct absorption spectra, but when mixed in the right proportions, the original didymium spectrum is reproduced. The atomic weights of the two elements differ but slightly from one another—praseodymium, 140.6; neodymium, 144.3. This reminds us of the resemblance between cobalt 58.97, and nickel 58.68. The history of cerium is rather curious. What was called cerium in 1803 was found to be a mixture of the elements lanthanum and cerium; and in 1841, lanthanum was found to be composed of two elements, didymium and lanthanum; the element didymium in 1879 was found to be a mixture of samarium and didymium; and in 1901, samarium was found to be a mixture of europium and samarium, while didymium, as just indicated, was found in 1885 to be a mixture of praseodymium and neodymium. There is quite a list of elements which have been announced but whose claims for recognition have not been generally accepted.

Meta-elements.—In 1887 W. Crookes argued very ingeniously that the closeness of the relationship between the metals of the rare earths makes it probable that these elements are modifications of one common element. For example, by a laborious process of fractionation, Crookes subdivided yttria into some eight components with different phosphorescent spectra, but the difference in the *chemical properties* of the fractions was so slight that, if it were not for a slight difference in the solubilities of the different fractions in ammonia, the fractions could not have been separated from one another. Crookes points out that the original yttrium passes muster as an element. It has a definite atomic weight, it enters into combination with other elements, and it can be separated from them as a whole. But the searching process of fractionation sorts the atoms of yttrium into groups with different phosphorescent spectra, and presumably different atomic weights, though from the usual chemical point of view all the groups behave alike. "Here, then, is a so-called element whose spectrum does not emanate equally from all its atoms; but some atoms furnish some, other atoms others, of the lines and bands of the compound spectrum of the element. Hence the atoms of this element differ probably in weight, and certainly in the internal motions they undergo." Assuming that the principle is of general application to all the elements, and is limited by our knowledge of tests delicate enough to recognize the simpler constituent groups of the different elements, it is inferred that there are definite differences in the internal motions of the several groups of the atoms of a chemical element. The seven series of bands in the absorption spectrum of iodine, for instance, may prove not to emanate from every molecule, but "some of these molecules may emit some of the series, others others, and in the jumble of all these molecules, to which is given the name 'iodine vapour,' the whole seven series are contributors." Ordinary elements thus represent the limits of the analysis of matter by chemical methods, while the spectra of some elements show that although chemically homogeneous they are probably mixtures of a number of distinct components. Crookes thus introduced the conception of what he called **meta-elements** for those fractional parts of an element which resemble one another so closely that the mixture of elements appears chemically homogeneous. The properties of the ordinary elements are supposed to be an average of the properties of aggregates of several meta-elements. The spectra of the meta-elements seem to furnish a decisive means of discriminating between these bodies which are otherwise similar in chemical properties.

Questions.

1. Indicate the points of similarity and contrast between the dioxides of barium, lead, and manganese.—*St. Andrews Univ.*
2. The molecular weight of litharge (an oxide of lead) is 223.1. The per cent. of lead is 92.8. The specific heat of lead is 0.031. Calculate the *exact* atomic weight of lead.—*Princeton Univ., U.S.A.*
3. Give the names and formulæ of the oxides of lead, and describe all that can be observed when each of these oxides is heated in an open crucible. From 1 gram of one of the oxides of lead, 1.269 grams of lead sulphate can be obtained: which of the oxides is it? ($\text{Pb} = 207$, $\text{O} = 16$).—*Sheffield Univ.*
4. Starting with the metals tin and aluminium, show how you would make sodium stannate and sodium aluminate; and in what respect the two processes differ.—*Amherst Coll., U.S.A.*

5. What substances are formed when tin, iron, and zinc are severally dissolved in hydrochloric acid? Describe experiments to show that the solutions obtained from tin and iron are reducing agents, and explain why they differ in this respect from the zinc solution.—*Sheffield Univ.*

6. What would be produced if tin foil were introduced into solutions of the following salts: (1) silver nitrate, (2) lead acetate, (3) copper sulphate, (4) copper chloride, (5) stannous chloride, (6) alum, (7) ferric chloride? Give equations.—*London Univ.*

7. When iron stands in damp air it rusts. When lead and silver are exposed to air they become tarnished. Explain what has taken place in each case. Do these changes cause any alteration in the weight of the metals?—*Aberdeen Univ.*

8. Explain the modern processes adopted for the smelting of lead from galena, and in the desilverization and softening of the crude metal.—*Board of Educ.*

9. Enumerate the chief sources from which the following metals are obtained: lead, tin, copper, and mercury. Describe briefly the method by which any one of these is produced from its ores.—*Aberdeen Univ.*

10. Describe the extraction of lead from galena and of silver from argentiferous lead. How may the oxides of lead be procured from metallic lead?—*St. Andrews Univ.*

11. When hydrogen is made in the ordinary way, to what is the disagreeable odour due? If tin were used instead of zinc, would you expect the same odour? (Explain from the method of purifying these metals).—*Amherst Coll., U.S.A.*

12. Explain how each of the following oxides can be prepared: Barium dioxide, sodium peroxide, nickelic oxide, lead peroxide. What is the action of hydrochloric acid on each of them, and by what chemical behaviour can they be classified into two groups?—*London Univ.*

13. What is meant by the valency of an element? How is it determined? Is it always the same for the same element? If not, have any relationships been found amongst the valencies of an element? What are the valencies of nitrogen, carbon, lead, iron?—*London Univ.*

14. To determine the equivalent of lead, Stas converted 103 grams of the pure metal into 164.775 grams of pure dry nitrate by evaporation with nitric acid. Assuming the atomic weight of oxygen to be 16, and of nitrogen 14.04, calculate the equivalent of lead. The specific heat of lead at 15° was found by Naccari to be 0.03. Calculate the atomic weight of lead and explain why a determination of the specific heats leads to a knowledge of the atomic weight.—*London Univ.*

15. Select any two substances out of the following list which could exist together in aqueous solution without any precipitation taking place:— $\text{Al}_2(\text{SO}_4)_3$, CuSO_4 , KNO_3 , ZnSO_4 , NaCl , $\text{Pb}(\text{NO}_3)_2$, FeCl_3 , $(\text{NH}_4)_2\text{CO}_3$, BaCl_2 , K_2S , $\text{Ca}(\text{NO}_3)_2$, NH_4Cl , KClO_4 .—*R. Galloway.*

16. Describe what occurs when hydrogen sulphide is passed (until there is no further action) into solutions of (a) caustic potash, (b) nitric acid, (c) stannic chloride, (d) ferric chloride, (e) chromic acid and sulphuric acid, (f) iodic acid, (g) sulphurous acid.—*Owens Coll.*

17. Give a short account of the sources and varieties of water used for drinking purposes. What potable waters are specially liable to give rise to "lead poisoning"?—*Calcutta Univ.*

18. Write an account of the oxides of lead, and show how the higher oxides may be formed from and converted into the lower.—*London Univ.*

19. One gram of copperas was dissolved in water and the solution made up to 250 c.c.; 50 c.c. was then mixed with ammonia and hydrogen peroxide and boiled for 5 minutes to expel the excess of hydrogen peroxide; the precipitated ferric hydroxide was dissolved in hydrochloric acid and titrated with a solution of titanous chloride each c.c. of which was equivalent to 0.001823 gram of Fe. An excess of potassium thiocyanate was used as inside indicator. 22 c.c. of the titanous chloride solution were needed to discharge the red colour of the ferric thiocyanate. Required the purity of the sample of copperas.—*E. Knecht and E. Hibbert.* Answer: 96.73 per cent. of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$.

20. What would be the effect produced if zinc-foil were introduced into solutions of the following salts: (1) silver nitrate, (2) lead nitrate, (3) copper sulphate, (4) copper chloride, (5) stannous chloride, (6) alum, (7) ferric chloride? Give equations.—*London Univ.*

21. Starting from lead sulphide, describe how you would prepare (a) lead, (b) litharge, (c) "red lead," (d) "white lead." How would you detect the

presence of lead peroxide and of lead chloride in a mixture of these two substances ?
—*Board of Educ.*

22. Thin layers of oxides form over the surfaces of many metals which have been exposed to the air. This prevents the metals being soldered together. The plumber makes firm joints by using resin or acid when he is soldering. Explain the action of these agents.

23. You are required to ascertain the atomic weight of tin from observations on the metal and on its chlorides. Assuming that you are supplied with these substances in a state of purity, describe in detail the nature of the experiments you would make.—*Punjab Univ.*

CHAPTER XLII

THE CLASSIFICATION OF THE ELEMENTS

§ 1. The Law of Octaves—J. A. R. Newlands.

The primary object of classification is to arrange the facts so that we can acquire the greatest possible command over them with the least possible effort.

The classification of the elements has long been an attractive subject. The elements have been classed into metals and non-metals; into acidic and basic, or, what amounts to the same thing, into electronegative and electropositive elements; they have been classed according to their valency; and also according to many other properties. In all these systems an element appeared in more than one class; or elements with but few properties in common were grouped together. The properties of the elements used as the basis of classification may also vary with the conditions under which the properties are observed.

In all chemical changes one property at least remains unaltered, and the more successful systems of classification have been based on this property, the atomic weights of the elements. The early efforts in this direction were seriously hampered by the uncertainty in the numerical values of the atomic weights. But after chemists had cleared up the confusion associated with the atomic theory left by Dalton, and obtained a consistent system of atomic weights, the results were more promising. True enough, between 1816 and 1829, J. W. Döbereiner noticed some regularities in the atomic weights of certain related elements, for he found that most of the chemically related elements either exhibited almost the same atomic weight—*e.g.* iron, cobalt, and nickel—or else exhibited a constant difference when arranged in sets of three. Thus, selecting one set from Döbereiner's list, and rounding off the modern atomic weights,

	Calcium.	Strontium.	Barium.
Atomic Weight	40	87	137
Difference		47	50

Many felt intuitively that the list of Döbereiner's triads was but a fragment of a more general law. Between 1863 and 1866, J. A. R. Newlands published a series of papers in which he arranged the elements in the ascending order of their atomic weights, and noticed that every succeeding eighth element was "a kind of repetition of the first." Thus,

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe
.

"In other words," said Newlands, "members of the same group of elements stand to each other in the same relation as the extremities of one or more octaves in music. This peculiar relationship I propose to provisionally term the law of octaves." Newlands noticed that elements belonging to the same group "usually" appeared in the same column, and he declared that all the numerical relations which had been observed among the atomic weights "including the well-known triads, are merely arithmetical results flowing from the existence of the law of octaves."

The "law of octaves" did not attract much attention, probably because faulty atomic weights seriously interfered with the arrangement.¹ Similar remarks apply to some papers by A. E. B. de Chancourtois in 1862, where it was proposed to classify the elements by their atomic weights. Chancourtois arranged the elements in a spiral line according to their atomic weights, and stated: *les propriétés des corps sont les propriétés des nombres*. Chancourtois' ideas were so much entangled with extraneous verbiage, and the truth was so much obscured by useless and faulty speculations, that his work lay buried for nearly thirty years, and it was only resurrected after Mendeléeff's system had become famous.

§ 2. The Periodic Law—D. I. Mendeléeff and L. Meyer.

The periodic series is a brilliant and adequate means of producing an easily surveyed system of facts which by gradually becoming complete will take the place of an assemblage of the known facts.—E. MACH.

D. I. Mendeléeff and L. Meyer, quite independently and, so far as we can tell, quite in ignorance of Newlands' and Chancourtois' work, obtained a far clearer vision of the "law of octaves" about 1869. Mendeléeff said: "When I arranged the elements according to the magnitude of their atomic weights, beginning with the smallest, it became evident that there exists a kind of periodicity in their properties." Otherwise expressed, if the elements be arranged in the order of increasing atomic weights, their *properties* vary from member to member in a definite way, but return more or less nearly to the same value at fixed points in the series. Mendeléeff continued: "I designate by the name 'periodic law' the mutual relations between the properties of the elements and their atomic weights, these relations are applicable to all the elements, and have the nature of a periodic ² function." Expressed more concisely, Mendeléeff's periodic law: The properties of the elements are a periodic function of their atomic weights. The early tables were very imperfect for the reasons stated above. Mendeléeff's tables of the atomic weights were designed to tabulate the elements in such a way as to exhibit the greatest number of relationships. The original tables were afterwards amended and modified owing to improved data and the discovery of new elements. The symbols of the elements with their atomic weights have been arranged on a helix, on a spiral, and in numerous other ways. Table LXIV., not

¹ When Mr. Newlands read a paper on "The Law of Octaves" at a meeting of the London Chemical Society in 1866, Prof. G. C. Foster said that any arrangement of the elements would present occasional coincidences, and inquired if Mr. Newlands had ever examined the elements according to their initial letters. Twenty-one years later the Royal Society awarded Newlands the Davy Medal for his discovery.

² A periodic function is one whose value repeats itself at regular intervals. The interval is called a "period."

TABLE LXIV.—THE PERIODIC SERIES OF THE ELEMENTS.

	Group 0	Group I R ₂ O	Group II RO	Group III R ₂ O ₃	Group IV RH ₄ R ₂ O ₂	Group V RH ₅ R ₂ O ₅	Group VI RH ₂ RO ₃	Group VII RH R ₂ O ₇	Group VIII Transition elements RO ₄
1		H 1-008							
2	He 3.99	Li 6.94	Be 9.1	B 11.0	C 12.00	N 14.01	O 16.00	F 19.0	
3	Ne 20.2	Na 23.00	Mg 24.32	Al 27.1	Si 28.3	P 31.04	S 32.07	Cl 35.46	
4	A 39.88	K 39.1	Ca 40.07	Sc 44.1	Ti 48.1	V 51.0	Cr 52.0	Mn 54.93	Fe 55.84; Co 58.97; Ni 58.68
5		63.57 Cu	65.37 Zn	69.9 Ga	72.5 Ge	74.96 As	79.2 Se	79.92 Br	
6	Kr 82.92	Rb 85.45	Sr 87.63	Yt 89.0	Zr 90.6	Nb 93.5	Mo 96.0	—	Ru 101.7; Rh 102.9; Pd 106.7
7		107.88 Ag	112.40 Cd	114.8 In	119.0 Sn	120.2 Sb	127.5 Te	126.92 I	
8	Xe 130.2	Cs 132.81	Ba 137.37	La 139.0	Ce 140.25	—	—	—	Os 190.9; Ir 193.1; Pt 195.2
9		—	—	—	—	—	—	—	
10	—	—	—	—	—	Ta 181.0	W 184.0	—	
11		197.2 Au	200.6 Hg	204.0 Tl	207.10 Pb	208.0 Bi	—	—	
12	Nt 222.5	—	Ra 226.4	—	Th 232.4	—	U 238.5	—	

very different in style from one of Mendeléeff's first tables, is one of the simplest modes of arrangement, perhaps the best.¹

The nine vertical columns are usually styled groups, and the thirteen horizontal lines series or periods. We have already studied the properties of the elements and of their compounds from the point of view of this system of classification. Our brief reviews of the families or groups will therefore suffice to emphasize the relationships of the members of any given group. Each short period, it will be observed, contains eight elements; and each long period either contains eighteen elements of which three are the so-called transition elements, or else it has provision made for eighteen elements. Hyphens are inserted in the spaces where the corresponding element is unknown. The elements in the first short series are sometimes called group elements or bridge elements, since they show a notable gradation of properties from one to the other, and serve as links or bridges between the different groups. The members of the next short period or series 3, are called typical elements because they have the typical properties and characteristics of the group, and show a rather wide divergence from neighbouring groups. Each typical element of the different groups diverges into two sub-groups.

The transitional elements.—It will be noticed that there is a distinct difference between the members of the odd and the even series. The alternate even series, say the fourth and sixth, resemble one another more closely than the successive members of the odd series, say the fifth and seventh. The lower oxides of the last members of the even series resemble in many ways the first members of the odd series. Thus the basic oxides of chromium and manganese are in many ways similar to the oxides of copper and zinc. Again, there are marked differences between the last members of the odd series (halogens) and the first members of the next even series (alkali metals). Those elements which cannot be placed in short periods fall in better with the last members of the even series and the first members of the odd series. Thus, iron, cobalt, and nickel fall between manganese and copper both with respect to chemical properties and atomic weights:

	... Cr	Mn	Fe	Co	Ni	Cu	Zn ...
Atomic weight	52	54.9	55.8	59.0	58.7	63.6	65.4
Specific gravity	6.9	7.4	7.8	8.7	8.8	8.9	6.9
Atomic volume	7.5	7.4	7.1	6.8	6.7	7.1	9.5


so also Ru—Rh—Pd → Ag come just after the sixth series, and Os—Ir—Pt → Au after the tenth series. The inert gases are considered to form a kind of transition between the last members of the odd series (halogens) and the first members of the even series (alkali metals), and consequently also, they only occur in the horizontal rows where transitional elements in the eight groups are absent.

The following arrangement, modified from one by T. Bayley (1882), emphasizes the relationship and yet the individuality of the subgroups, the character of the transition elements, etc.²

¹ Other schemes have advantages as well as disadvantages over Mendeléeff's arrangement, so that there is no special reason why any one should be here given preference.

² The meaning of "Protyle" and of "Extinct Elements" will be indicated in the discussion later on.

Protyle.

																		H																									
																		He—Li—Be—B—C—N—O—F																									
																		Ne—Na—Mg—Al—Si—P—S—Cl																									
																																											
A	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br																										
Kr	Rb	Sr	Y	Zr	Nb	Mo	—	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	—																										
Xe	Cs	Ba	La	Ce	—	—	—	—	—	—	—	—	—	—	—	—	—																										
—	—	—	—	—	Ta	W	—	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	—	—																										
Nt	—	Ra	—	Th	—	U	—	—	—	—	—	—	—	—	—	—	—																										
																		<i>Extinct Elements.</i>																									

This table also emphasizes the fact that while the atomic weights of the elements progressively increase, their properties recur at definite intervals; no well-known elements are omitted from the scheme; with three exceptions the order is that of the atomic weights, and the elements fall into virtually the same groups as would have been obtained had they been grouped according to their chemical behaviour.

Most of the well-defined physical and chemical properties of the elements are periodic—valency, specific gravity, atomic volume, melting point, hardness, malleability, ductility, compressibility, coefficient of expansion, thermal conductivity, latent heat of fusion, refraction equivalents for light, colour, electrical conductivity, magnetic power, etc. When the numerical values of these properties and the atomic weights of the elements are tabulated on squared paper, a curve is obtained which is broken up into periods, as we have found to be the case in Fig. 151. The properties of analogous compounds of the elements—molecular volumes, melting points, boiling points, stability of the hydroxides, colour, etc.—are very often periodic. The heats of formation of the chlorides, Fig. 114, is a good illustration. The specific heats of the elements are unique in furnishing a non-periodic curve. According to Dulong and Petit's law, if x denotes the specific heat of an element with an atomic weight, y , we have $xy = 6.4$. This is obviously a hyperbolic not a periodic curve like Fig. 151.

§ 3. The Gaps in Mendeléeff's Tables of the Elements.

The periodic law has given to chemistry that prophetic power long regarded as the peculiar dignity of its sister science, astronomy.—H. C. BOLTON.

Both Meyer and Mendeléeff considered it necessary to leave gaps in their tables for undiscovered elements, and more particularly in order to keep certain related elements in the same vertical column. Mendeléeff boldly prophesied that the missing elements would be discovered later, and in some cases even predicted their properties in considerable detail.

For instance, when Mendeléeff announced the law, there were two blank spaces in group III., the missing elements were called *eka-aluminium* and *eka-boron* respectively; and another space below titanium in group IV., the missing element in this case was called *eka-silicon*. The hypothetical character of these elements was considered to be an inherent weakness of the law, but the weakness was turned to strength when gallium, scandium, and germanium subsequently appeared duly clothed with those very properties which fitted closely with Mendeléeff's audacious prognostications. This hit attracted considerable attention, and served to strengthen the faith of chemists in the fundamental truth of the periodic law. In illustration the case of *eka-silicon* and germanium is quoted side by side in Table LXV.

TABLE LXV.—COMPARISON OF PREDICATED AND OBSERVED PROPERTIES OF GERMANIUM.

Eka-silicon, Es (predicted in 1871).	Germanium, Ge (discovered in 1886).
<p>Atomic weight, 72 Specific gravity, 5.5 Atomic volume, 13 Element will be dirty grey, and on calcination will give a white powder of EsO_2 Element will decompose steam with difficulty Acids will have a slight action, alkalies no pronounced action</p>	<p>Atomic weight, 72.3. Specific gravity, 5.47. Atomic volume, 13.2. The element is greyish-white and on ignition furnishes a white oxide GeO_2. The element does not decompose water.</p>
<p>The action of sodium on EsO_2 or on EsK_2F_6 will give the element</p>	<p>The element is not attacked by hydrochloric acid, but it is attacked by aqua regia. Solutions of KOH have no action, but the element is oxidized by fused KOH. Germanium is made by the reduction of GeO_2 with carbon, or of GeK_2F_6 with sodium.</p>
<p>The oxide EsO_2 will be refractory and have a sp.gr. 4.7. The basic properties of the oxide will be less marked than TiO_2 and SnO_2, but greater than SiO_2.</p>	<p>The oxide GeO_2 is refractory and has a sp.gr. 4.703. The basicity is very feeble.</p>
<p>Eka-silicon will form a hydroxide soluble in acids, and the solutions will readily decompose forming a metahydrate</p>	<p>Acids do not precipitate the hydrate from dilute alkaline solutions, but from concentrated solutions, acids precipitate GeO_2 or a metahydrate.</p>
<p>The chloride EsCl_4 will be a liquid with a boiling point under 100° and a sp.gr. of 1.9 at 0°</p>	<p>Germanium chloride, GeCl_4, boils at 86°, and has a sp.gr. at 18°, 1.887.</p>
<p>The fluoride EsF_4 will not be gaseous</p>	<p>The fluoride $\text{GeF}_4 \cdot 3\text{H}_2\text{O}$ is a white solid mass.</p>
<p>Eka-silicon will form a metallo-organic compound $\text{Es}(\text{C}_2\text{H}_5)_4$, boiling at 160°, and with a sp.gr. of 0.96.</p>	<p>Germanium forms $\text{Ge}(\text{C}_2\text{H}_5)_4$, which boils at 160°, and has a specific gravity slightly less than water.</p>

The confirmations of Mendeléeff's predictions of the properties of *eka-aluminium* (gallium) and of *eka-boron* (scandium) were equally striking. This dramatic achievement focused attention on the generalization; but it is only fair to say that the predictions and their subsequent verification are not such positive proofs of the truth of the periodic law as some suppose.

It is certainly wrong to say, as C. Winkler did, "it would be impossible to imagine a more striking *proof* of the doctrine of periodicity of the elements than that afforded by this embodiment of the hitherto hypothetical eka-silicon," because gaps appeared in some of the older systems of classification, and the properties of the missing members could have been predicted, and the atomic weights estimated by analogy with the other members of the family, quite independently of, and in some cases better than, the periodic law.

§ 4. The Applications of the Periodic Law.

A natural law only acquires scientific importance when it yields practical results, that is, when it leads to logical conclusions which elucidate phenomena hitherto unexplained, when it directs occurrences till then unknown, and especially when it calls forth predictions which may be verified by experiment.—D. I. MENDELÉEFF.

Mendeléeff pointed out that the periodic law could be employed in:

1. The classification of the elements;
2. The estimation of the atomic weights of elements not fully investigated;
3. The prediction of the properties of hitherto unknown elements;
- and 4. The correction of atomic weights.

1. The classification of the elements.—T. H. Huxley (1864) has said: "By the classification of any series of objects, is meant the actual or ideal arrangement together of those which are like, and the separation of those which are unlike; the purpose of this arrangement being to facilitate the operations of the mind in clearly conceiving and retaining in the memory the characters of the objects in question." The periodic system is undoubtedly superior to all the older methods of classification, for the law makes it possible to build up a system of the greatest possible completeness, free from much arbitrariness, and it furnishes strong circumstantial evidence of the correctness of the reasoning employed by Cannizzaro to deduce values for the atomic weights of the elements.

2. The estimation of the atomic weights of the elements.—On account of practical difficulties, it is not always possible to fix the atomic weights of some elements by vapour density determinations (Avogadro's rule), and by specific heat determinations (Dulong and Petit's rule), and the atomic weights of these elements were frequently assigned on somewhat uncertain grounds. According to C. L. Winkler, indium has the equivalent weight 37·8. The correct atomic weight must be some multiple of this, and for no special reason, the atomic weight was once taken to be $37·8 \times 2 = 75·6$. In that case, indium would fall between arsenic and selenium, where it would be quite mis-matched. Mendeléeff proposed to make indium trivalent, like aluminium, so that the atomic weight became $37·8 \times 3 = 113·4$, and the element fell in the table between cadmium and tin where it fits very well. The subsequent determination of the specific heat of indium, 0·0577, corroborated the change made by Mendeléeff in the atomic weight from 75·6 to 113·4. Beryllium, uranium, and a number of the rare earths at one time did not fit very well into the table, but Mendeléeff's alteration of the supposed atomic weights to make these elements fit the table were subsequently justified by vapour density determinations of the volatile chlorides, or by specific heat determinations.

3. The prediction of the properties of hitherto undiscovered elements.—In order to avoid introducing new names when speaking of unknown elements, Mendeléeff designated them by prefixing a Sanscrit numeral—eka (one), dwi (two), tri (three), etc.—to the names of the next lower analogous elements of the odd or even numbered series of the same group. Thus, the unknown elements of group I. will be called eka-cæsium and dwi-cæsium. Were strontium unknown, it would be called eka-calcium. In addition to the prediction of germanium, gallium, and scandium already discussed, Mendeléeff foretold the possible discovery of eka- and dwi-cæsium; eka-niobium— $\text{En} = 146$; of eka-tantalum— $\text{Et} = 235$; of dwi-tellurium— $\text{Dt} = 212$; and of the analogues of manganese: eka-manganese— $\text{Em} = 100$; and tri-manganese— $\text{Tm} = 190$.

The case of the so-called inert gases is of more recent date. The discovery of argon and helium could not have been predicted from Mendeléeff's periodic law, but after these elements had been discovered, accommodated in the periodic table between the strongly acid halogen family and the strongly basic alkali metals, the probable existence of other similar inert gases was indicated. When an exhaustive search was made krypton, neon, and xenon were discovered with properties and atomic weights which could have been predicted from the arrangement made for argon and helium in Mendeléeff's table.

4. The correction of the values of atomic weights.—If the atomic weight of an element does not fit with the regular course of, say, the atomic volume curve, Fig. 151, the atomic weight is probably in error. Thus, the atomic weights of platinum, iridium, and osmium at that time were probably too high, and subsequent determinations verified this inference. Thus the atomic weights of these elements were:

	Platinum.	Iridium.	Osmium.
In 1870	196.7	196.7	198.6
In 1912	195.2	193.1	190.9

There are also some misfits in the table as we have it to-day, owing to the fact that at least three pairs of elements would be mis-matched if they were simply classed according to their atomic weights: argon (39.88) and potassium (39.10); cobalt (58.97) and nickel (58.68); and tellurium (127.5) and iodine (126.92). The case of iodine and tellurium has been studied very closely. Iodine most certainly belongs to the same group as the other halogens, and tellurium to the selenium group, and these elements are accordingly placed in these groups in spite of the fact that if their atomic weights were alone considered tellurium would be ranked with the halogens, and iodine with selenium. B. Brauner supposes that ordinary tellurium is a complex containing α - and β -tellurium; but tellurium has been melted, sublimed, oxidized, hydrogenized, phenylated, dissolved, crystallized, and precipitated; yet nothing but failure has followed all attempts to get an atomic weight lower than iodine. Hence in spite of the fact that "the laws of nature admit of no exception" (p. 102), faith in the law has led chemists to allocate these discordant elements according to their chemical properties and not according to their atomic weights. This method must be dubbed "unscientific,"¹ but the

¹ Bode's law of astronomy successfully predicted the asteroids and allocated their proper place in the solar system; but the subsequent discovery of Neptune did not agree with Bode's law. The "law" was accordingly abandoned and it is now regarded as a curiosity.

circumstantial evidence justifies the procedure, in the expectation that a consistent system will ultimately grow from the truth and error engrafted into the "law." It is not very probable that the principle underlying the periodic law will be abandoned because it is founded on a vast assemblage of facts of different kinds; and because it seems to be plastic enough to fulfil subsequent requirements.

§ 5. Some Defects in the Periodic Law.

The allocation of hydrogen in the table, as already indicated, is difficult, because hydrogen seems to be without companions. It is univalent, and thus falls either with the alkali metals (D. I. Mendeléeff) or with the halogens (O. Masson). Hydrogen is electropositive like the alkali metals, but it is certainly not now considered to be a metal. It can be displaced by the halogens from the organic compounds, and it forms hydrides with some of the metals not at all unlike the halogen salts. Thus the position of hydrogen has not been definitely settled. Indeed, *hydrogen appears to be a "rogue" element quite out of place in the general scheme.* Some suppose that hydrogen is a member of a series of independent elements yet unknown; the supposed last member of the series is called "proto-fluorine"; so also the elements "proto-beryllium" and "proto-boron" have been invented, the former with an atomic weight 1.33, and the latter, 2. All this, however, is mere speculation.

There is a difficulty about the allocation of the rare earths. Most of them could be distributed about the table according to their atomic weights, or they could be relegated to a class by themselves. B. Brauner, who has made a special study of the rare earths, considers that they should all be grouped together with cerium so that "Ce, 140.25" in the table stands for:

Ce, 140.25; Pr, 140.6; Nd, 144.3; Sa, 150.4; Eu, 152; Gd, 157.3; Tb, 159.2; Dy, 162.5; Er, 167.7; Tm, 168.5; Yb, 172.0 . . .

This has been called the asteroid theory of the rare earths. The properties of the rare earths, however, are not well enough known to give us much confidence in the various proposals which have been made for dealing with them; and consequently, Mendeléeff considered that the installation of these elements should be deferred. The better known elements of the rare earths fit well enough into the table, thus, ytterbium—Yb, 172—fits into group III., series 10, etc.

If the properties of the elements are dependent on their atomic weights the existence of two elements with different properties and approximately the same atomic weights should be impossible. Hence the difficulty with elements like cobalt and nickel; ruthenium and rhodium, etc. The peculiarities of these elements would never have been suspected from the periodic law.

Some elements are allocated places in the table according to their atomic weights in opposition to their properties. For instance, copper, silver, and gold fall into one group with the alkali metals. The trivalency of gold appears to be unconformable with the valency of its companions, although in its present position the series: PtCl_4 , AuCl_3 , HgCl_2 , and TlCl is suggestive. Thallium is very like lead, but its sulphate and some other salts are quite different from lead salts. At least three pairs of elements have been placed according to their properties irrespective of

their atomic weights, as indicated by the "misfits" mentioned in the preceding section.

Some elements which appear to be chemically similar are separated in the table. For example, copper and mercury; silver and thallium; barium and lead; etc. The position of these elements in the table gives no hint of these characters. Still, it might be argued that these elements exhibit many essential differences. Thus the physical properties of mercury and copper, and the properties of the cupric and mercuric chlorides and sulphates show great contrasts. The stability of cuprous and mercurous chlorides is also very different. Lead and barium peroxides appear to have a different constitution. The unstable thallium sesquioxide, Tl_2O_3 , corresponds with the other—more stable—sesquioxides in the group, but there are many important points of resemblance between thallium and the alkali metals, and between silver and lead.

G. Wyruboff (1896) has the idea that the periodic system is "a very interesting and highly ingenious table of the analogies and the dissimilarities of the simple bodies—a mere *catalogue raisonné* of the elements"; and further, "since the laws of nature admit of no exception, the periodic law must be considered as a law of nature definitely established which must be accepted or rejected as a whole." Wyruboff's proposal to reject the periodic law is somewhat precipitate, for we do not feel quite satisfied that the supposed misfits are not due to defective knowledge. There is a distinction between "failure" and "incomplete success."

§ 6. The Unitary Theory of Matter.

Chemistry marches towards its goal and towards its perfection by dividing, subdividing, and subdividing still again.—A. L. LAVOISIER, 1789.

Simplicity is the seal of truth. Nature is wonderfully simple, and the characteristic mark of a childlike simplicity is stamped upon all that is true and noble in nature.—M. SENDIVOGIUS (1650).

Belief in the simplicity of nature is not logic but faith pure and simple. It is one of those insidious and dangerous *tacit* assumptions which often creep into scientific theories. Tacit assumptions are "dangerous" because they are usually made unconsciously, so that they appear to be self-evident truths, and prevent our harbouring the shadow of a doubt of their insidious character. True enough, as H. Poincaré has said: "Every generalization supposes in a certain measure a belief in the simplicity of nature . . . every law is considered to be simple until the contrary has been proved," but faith in this dogma has sometimes led men astray. In consequence, hypotheses have frequently flourished in spite of experimental evidence to the contrary. It is not at all uncommon to find a "law" which appears to be quite simple when the methods of measurement are crude and approximate, but becomes exceedingly complex when more accurate data are available. We have found this to be the case, for instance, with Boyle's law, Charles' law, etc. Another example will now be given.

There was a marked tendency among the earlier Greek philosophers to postulate one single kind of matter a *πρώτη ἄλη*—(prōtē first, (ylē) matter—or primal element. This *prima materia* ("potential matter") was supposed to consist of parts which when grouped in different ways produced the various kinds of matter considered by them to be elemental.

The hypothesis is sometimes called the unitary theory of matter: all the different forms of matter in the universe are derived from one and the same primordial element.

The philosopher Anaximenes regarded air as the primal element; Herakleitos, fire; Pherekides, earth; Prout, hydrogen; and modern chemical philosophers, electrons, and possibly ether. Thales of Miletus, who flourished in the sixth century B.C., considered that water was the first principle. Thales' doctrine had many supporters—J. B. van Helmont (1682), etc.—it lived for nearly twenty-five centuries; and it was demolished in 1770 when Lavoisier demonstrated that water cannot be changed into earth (p. 39).

In 1815 and 1816 W. Prout tried to show that the atomic weights of the elements were exact multiples of the atomic weight of hydrogen, so that an atom of any element must weigh a certain number of times as much as an atom of hydrogen. Hence added Prout, "we may almost consider the $\pi\rho\omega\tau\eta$ $\psi\lambda\gamma$ of the ancients to be realized in hydrogen." According to Prout's hypothesis, the elements are different aggregates of the atoms of primordial hydrogen; that is, the different elements are polymers of hydrogen; in consequence, within the limits of experimental error, the atomic weights of the different elements should be expressible by whole numbers when the atomic weight of hydrogen is unity.

Many writers, attracted by its apparent simplicity, gave unqualified support to Prout's hypothesis; but an impartial review of the facts, with very much more refined data than were available in Prout's day, led J. S. Stas (1860-1865) to state: "I have arrived at the absolute conviction, the complete certainty, so far as is possible for a human being to attain to certainty in such matters, that the law of Prout is nothing but an illusion, a mere speculation definitely contradicted by experience." Many have tried to reconcile Prout's hypothesis with facts by changing the standard of reference to an imaginary primordial element with an atomic weight equivalent to half an atom of hydrogen (C. Marignac, 1844); then to a quarter of an atom (J. B. Dumas, 1859). But having once begun to divide the hydrogen atom, there was no limit to the process of subdivision, and the hypothesis could then be made to fit any conceivable set of atomic weights. This tinkering with Prout's hypothesis brought it, for a time, into disfavour.

The elements, as we have seen, are substances which have never been resolved into simpler substances; we did not say that the elements could never be decomposed into a more primitive form (or forms) of matter. The atom of an element can be defined as a substance whose parts are held together by a force superior to any which has yet been brought to bear upon it. The elements have hitherto proved undecomposable, but it is quite conceivable that they are not absolutely undecomposable. Much circumstantial evidence has accumulated in recent years which makes it increasingly difficult to believe that Dalton's atoms are absolutely indivisible; and to deny that all the different elements have been formed from one homogeneous simple primal form of matter which G. Hinrichs has called pantogen (1857); W. Crookes, protyle (1886); and J. L. G. Meinecke, urstoff (1817). It has been said that protyle "is matter generalized, stripped of its distinctions, the same from whatever source derived; it is matter in potency rather than in act; intangible, inacces-

sible to sense perception, probably indifferent to the solicitations of gravity."

The idea that all the different forms of matter represent different stages in the growing complexity of one single elemental form of matter, has haunted the human mind from the earliest times, and the belief that "unity is the law of God," or that *simplex veri sigillum*, has proved so peculiarly fascinating that a remarkable number of speculations have been spun about this theme. These ingenious and subtle systems are admirable works of art, but they are outside the realm of science so long as the evidence on which they are founded remains intangible and vague. In recent years still another hypothesis is claimed to have been woven with circumstantial evidence which is rather more substantial than the incomprehensible fictions of the metaphysicians, for it can be examined and tested by comparison with gross material facts. It is supposed that the elements have been developed by the condensation of a primitive form of matter; and, that the different elements, as W. Crookes expresses it, have been evolved by a kind of struggle for existence where elements, not in harmony with their environment, have either disappeared—*extinct elements*—or have never existed; where elements—*asteroidal elements*—have come into being and survived only on a limited scale as is the case with the *rarer elements*; and where other elements predominate because the surrounding conditions have been favourable to their formation and preservation, *e.g.* the *common elements*. This circumstantial evidence can now be outlined:

1. Why do so many atomic weights approximate to whole numbers?—The International Table of Atomic Weights, 1916, contains 83 elements, and of these, 43 are integers within one-tenth of a unit. This fact has nothing to do with Prout's discarded hypothesis, but the numbers are facts which can hardly be due to chance, because the probability of this occurring is exceedingly small—something like one in 20,000 millions. It is argued that the approximation of the atomic weights of so many elements to whole numbers cannot be reasonably attributed to a fortuitous coincidence.

2. Why do so many groups of the elements exhibit family relationships?—If the elements are totally distinct and independent of one another, it is exceedingly difficult to reconcile the regularities and analogies of the elements in different groups revealed by a study of Mendeléeff's law, the main characteristic of which is *relationship*. The periodic law dimly foreshadows an identical origin or common parentage of families of elements. A study of the alkali metals, the metals of the alkaline earths, the halogens, etc., makes it highly probable that the different elements of one family, at least, have been formed by the conglomeration of monads or atomicules formed of the same primal matter so as to build up ordinary atoms of different sizes or shapes, that the evolution has progressed from homogeneity to heterogeneity. In other words, said C. R. A. Wright (1873), the so-called elements are allotropic modifications of a primitive matter, and they differ from one another in the amount of latent energy they contain per unit mass. The idea is illustrated by the homologous series of carbon compounds. For instance,

	Ethene.	Propene.	Butene.	Pentene.	Hexene.
	C_2H_4	C_3H_6	C_4H_8	C_5H_{10}	C_6H_{12}
Molecular weight	28	42	56	70	84

apparently by the polymerization of an increasing number of CH_2 groups. This series rises in a similar manner to nearly $\text{C}_{30}\text{H}_{60}$, with a regular increase of 14 in the molecular weight. Dobereiner's triads too have emphasized similar "constant" differences in the atomic weights of related elements. Hence, argued D. Carnelley (1885), if a body we know to be compound can play the part of an element, there is some plausibility in the suggestion that the elements themselves are not absolutely simple.

3. Why are closely related elements so often associated together in nature?—Chemists have long been struck with the peculiar way certain elements occur in the half-mile crust of the earth. Although no disturbing agent has been recognized at work in nature whereby the different kinds of elements are sorted like to like, yet certain groups of elements nearly always occur in juxtaposition. These individual elements are not plentifully distributed, and they are not easy to separate from one another—*pares cum paribus facillime congregantur*. For example, cobalt is perhaps never quite free from nickel, and *vice versa*; silver is almost invariably associated with lead ores and with gold; cadmium with zinc ores; selenium with the sulphur of pyrites; the members of the two groups of the platinum elements; and the rare earths. These associations and co-mixings cannot be entirely due to chance, for these elements are neither plentifully distributed nor have they any marked chemical affinity for one another. Consequently, it has been suggested that the elements in question were formed from some common material under almost identical conditions, and where slight variations in the conditions led to the almost simultaneous formation of closely related elements. Environment has determined the path of the evolution of the elements.

Additional circumstantial evidence for the unitary theory has been obtained from (4) the grouping of the spectral lines (p. 1003); (5) the magnetic perturbation of the spectral lines (p. 1005); (6) the phosphorescent spectra of the meta-elements (p. 987); (7) spectra of the stars and nebulae (p. 1006); (8) electric discharges in attenuated gases (p. 1010); and (9) radioactivity (pp. 1010 *et seq.*). This evidence will now be reviewed.

§ 7. Evidence for the Unitary Theory from Spectrum Analysis.

The final impression our mind receives on contemplating these fundamental relations is that of a wonderful mechanism of nature, the functions of which are performed with never-failing certainty, though the mind can only follow them with difficulty, and with a humiliating sense of the incompleteness of its perception.—J. J. BALMER.

According to modern theories, elaborated in text-books on physics, light is produced by periodic vibrations of the æther; and light waves in the æther can be inaugurated by the motions of very small particles—atoms, or infra-atoms—vibrating with a certain frequency.¹ Each particle of matter vibrating with the right frequency can set up the periodic vibrations in the æther which we call light. The eye only recognizes as light the vibrations of the æther ranging between the extreme red end of the visible spectrum (wave-length: 81 millionths cm.), and the extreme violet

¹ It is not at all unlikely that the absorption of actinic energy by say, insulated chlorine (p. 281), is attended by a change in the vibratory motion of the atoms in the chlorine molecules, and that the vibratory energy is transmuted into an increase in the translatory motion of the molecules which causes the rise of temperature and other phenomena indicated on p. 140.

end (wave-length: 36 millionths cm.). The different vibrations are "sorted" by a glass prism (p. 422) as spectral bands and lines or a complete spectrum. Each particular band or line is an effect of the periodic vibrations of the particles. The constancy of the spectral lines in the spectrum of a gas shows that the motions of the vibrating particles must be remarkably regular. The period of vibration of the particles, as revealed by the spectroscope, conditions a chemical property which in the opinion of R. Bunsen is as constant and fundamental as the atomic weight of the element, and which may be determined with almost astronomical exactitude; indeed, in 1900, the International Congress in Paris accepted a proposal by A. de Gramont that no new substance be recognized as an element until its spark spectrum has been measured, and shown to be different from that of every other known form of matter. The large number of spectral lines in the spectra of many elements, e.g. iron, shows that the vibrations which give rise to the spectrum must be remarkably complex. According to J. N. Lockyer, the lines are connected with the different vibratory powers of different portions of an atom. In other words, an atom is a complex aggregate of particles; and hence H. A. Rowland was led to say that a grand piano must be a simple mechanism in comparison with a vibrating atom. All this is supposed to imply that the spectral lines are due to the vibrations of particles smaller than the atom, and that the atom itself is a complex system made up of vibrating corpuscles (cf. p. 987). Indeed, *the atom itself is now considered to be a kind of miniature planetary system embracing hundreds of corpuscles each spinning in its own tiny orbit.* This recalls the speculation of p. 146, where the individual atoms were supposed to whirl rhythmically in minute orbits to form molecules, much as the planets in the solar system revolve about the sun. The idea is quite an old one. S. Brown, for example, in 1843, said:

The atom is by no means essentially or even potentially indivisible. There is a possibility and likelihood that within the atom is another nameless world of the universe. Under the sky of the atom proceeds worlds of material existence as different from atoms as atoms from compound particles; as compound particles from crystal shapes; as crystal shapes from stars and planets; as stars and planets from solar systems; or as solar systems from firmaments.

The homologous grouping of the spectral lines.—At first glance, the spectral lines of a given element seem to be so complex that any attempt to reduce them to order appears to be hopeless. The spectrum of iron, for instance, contains over 4000 lines. G. J. Stoney (1870) noticed a semblance of order in the spectrum of hydrogen, and J. J. Balmer (1885) found that the wave-lengths of most of these lines can be computed if whole numbers be substituted for m in a formula of the type $3647.2 m^2 / (m^2 - 4)$; and J. S. Ames (1890) showed that the error involved in the wave-lengths of the 29 known hydrogen lines is accurate to about one part in 100,000. Thus.

		H α	H β	H γ	H δ	H ϵ
	$m =$	3	4	5	6	7
Wave-lengths {	Calculated .	6564.97	4862.93	4342.00	4103.11	3971.40
	Observed .	6564.96	4862.93	4341.90	4103.10	3971.40

In the search for analogous laws in the spectra of other elements, it was found that the lines of many apparently chaotic spectra can be resolved

into a number of regular series superposed one on the other. Thus, the spectral lines of the elements can be arranged in groups or series. The spectrum of helium, for example, can be resolved into eight such series, each of which has lines which show a regular decrease in intensity with vibration frequency. Oxygen has a spectrum with six superposed sets of lines; hydrogen and the alkali metals have each four such series. The metals in question have monatomic molecules, and hence it inevitably follows that the atoms have a complex constitution, and that the spectral lines are not all produced by the vibrations of the same kind of particles; atoms are complex aggregates of vibrating corpuscles.

Although the spectrum of each element is characteristic of that element and of no other, yet the spectrum of each of the five alkali metals can be

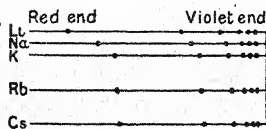


FIG. 333.—One Series of Homologous Spectra of the Alkali Metals Spaced according to their Atomic Weights (C. Runge, 1912).

resolved into four homologous series such that every line in one spectrum is represented by a corresponding line on all the other spectra; and generally, the spectra of related elements form natural or homologous groups. The corresponding lines do not necessarily all occur in the visible spectra, *e.g.* the dark red line of the rubidium spectrum corresponds with a line in the ultra-red in the other spectra; and the yellow lines of the sodium spectrum are homologous with lines in the ultra-violet spectra of the other elements. The spectra of related elements seem to be subject to a law of homology which is closely connected with their atomic weights. Each series of homologous lines appears to contract in passing from one element to another so that as the atomic weight increases, the lines at the red end of the spectrum appear to open further apart, and at the violet end, to draw closer together. Thus, by representing the spectral lines of the alkali metals by dots, Fig. 333, so that their limits at the violet end coincide, it will be seen that the lines of this particular series contract in passing from lithium to caesium, and the relation between the contractions and the atomic weights of the elements is shown in the diagram by making the distances between the lines represent the magnitude of their atomic weights.

There is thus a family likeness between the spectra of related elements, and a close connection has been traced between the atomic weights and the differences in the vibration frequencies of the lines in the members of a given family of elements. The atomic weights of groups of related elements with their related physical and chemical properties differ by fixed definite values, and the spectra of these elements show that the masses of their atoms affect their rates of vibration in a similar manner. This is taken to mean that not only is an atom of an element a complex composed of different aggregates of particles, but the atoms of allied elements are built up of similar aggregates of particles; and the observed differences in the spectra of allied elements results from differences in the way these aggregates are arranged in the atom.

The atomic weights of certain elements have been calculated from the relationship of the spectral lines of allied elements. For instance, the spectral lines of the magnesium-calcium family can be grouped in three

pairs; and the stronger lines of the radium spectrum are analogous with the stronger lines of the spectra of the other elements of the same family. The two lines of each of the three pairs are the same distance apart for any one element, but the distances apart of the lines in passing from one element to another increase in a regular manner with increasing atomic weights. Thus:

	Magnesium.	Calcium.	Strontium.	Barium.	Radium.
Atomic weight	24.36	40.1	87.6	137.4	?
Distance	91.7	223	801	1691	4855.6 units.

By plotting the atomic weights as functions of the distances between the lines, extrapolation furnished C. Runge and J. Precht (1903) with an approximate value for the atomic weight of radium. The extrapolation is easier if logarithms of the atomic weights and the distances apart of the lines be plotted. The graph for a group of chemically related elements then lies in a straight line, as indicated in Fig. 334. Lecoq de Boisbaudran (1886) evaluated the atomic weight of gallium and of germanium by a similar method.

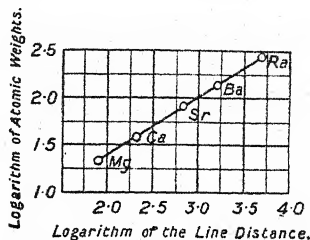


FIG. 334.—Relation between the Atomic Weights and the Distance apart of the Spectral Lines of the Mg-Ca Family of Elements.

The action of a magnetic field on the spectral lines—Zeeman's effect.—

P. Zeeman (1897) found that the yellow line characteristic of the spectrum of sodium is displaced or "perturbed" because it is split into two separate and distinct lines. And generally, the lines in the normal spectrum of an element are usually broadened into doublets, triplets, sextets, octets, or still more complex groups. The effect will be evident from Fig. 335. Different lines of the elementary spectra are perturbed in different ways—

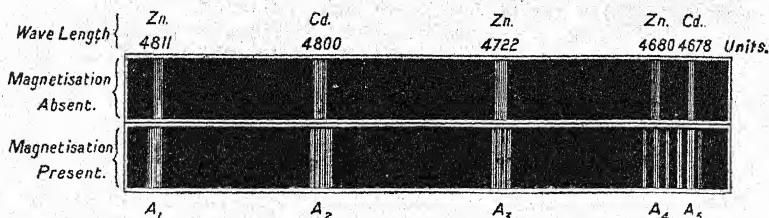


FIG. 335.—Effect of Magnetic Field on the Spectral Lines of Zinc and Cadmium—Zeeman's effect.

some lines, originally single, may remain single; others give rise to doublets, triplets, quartets, octets, etc. There is a certain pathetic interest attaching to this phenomenon because it is said that the last experiment made by M. Faraday (March 12, 1862) was an attempt to find if any spectral change could be detected when the source of the light is placed in a magnetic field. Faraday's experiment failed because the only available magnet was not strong enough. The perturbation of the spectral lines in a strong magnetic field is supposed to be due to alterations in the mode of vibrations of the sub-atomic particles

under a magnetic stress. The observed displacement depends on the strength of the magnetic field. If the spectrum be produced by the vibration of atoms of the same kind, it would be difficult to explain why the mode of vibration of atoms of the same kind can be altered in this way. It is more reasonable to suppose that the different effects produced by a magnetic field on the different lines of the spectrum of a substance are due to the presence of different kinds of particles in the incandescent element. This simply means that at the high temperature of the luminous object, the atoms of the element are dissociated into simpler "sub-atoms" or corpuscles which are probably electrified when they are called electrons. Owing to the fact that corresponding lines in the spectra of related elements—for instance, magnesium, zinc, cadmium—are similarly affected so that groups of lines in the different spectra undergo the same modification, it is inferred that the vibrating corpuscles which produce the A_1, A_2, A_3, \dots lines in the spectrum of zinc are the same as the vibrating corpuscles which produce the A_1, A_2, A_3, \dots lines in the spectrum of cadmium, etc. Hence, not only is an atom a complex association of different corpuscles, but the atoms of allied elements probably contain certain groups of corpuscles, or sub-atoms, in common. "These observations," said T. Preston (1899), "lend some support to the idea so long entertained merely as a speculation, that all the various kinds of matter, all the various so-called elements, may be built up in some way of the same fundamental substance."

The spectra of stars and nebulae.—The nature of the spectrum of an incandescent body depends upon the temperature (p. 425); and since the spectra of many elements have been observed at several different temperatures, it is possible to get a rough idea of the temperature of any incandescent element from the appearance of its spectrum. For instance, the flame spectrum of sodium has two yellow lines close together, while the spark spectrum has a pair of lines in the orange and in the green; the flame spectrum of thallium has a single green line, and the spark spectrum has in addition some bands in the violet. This fact is explained by the assumption that at the higher temperatures the atoms of an element are aggregates of sub-atoms, each of which has its own characteristic mode of vibration. The spectra of the nebulae, stars, and suns thus enable rough approximations of their temperature to be made.

The sun, stars, and nebulae, says A. M. Clerke, form so many celestial laboratories where the nature and mutual relations of the chemical elements may be tried by more stringent tests than sublunary conditions afford. In the very hottest stars¹ (estimated temperature $25,000^\circ$)—*e.g.* β -Crucis—comparatively few chemical elements can be detected, while in the cooler red stars—*e.g.* Betelgeuse—the number of spectral lines is comparatively large and a large number of elements are present. New elements appear to be introduced at each stage in the cooling of hot stars, so that elements

¹ In the orthodox works on astronomy, we are told that the nebulae are the material out of which stars are made, and that in their forms, aggregations and condensations, the process of evolution of stars and suns can be traced. The nebulae begin either as hot attenuated gases or as clouds of cold meteoric stones which gradually gain heat as they clash together until finally they are converted into gases, and then gradually cool by radiation. Hence the cooler stars are either young or old, heating up or cooling down. The hottest stars are in their prime.

which were non-existent in the hotter stars make their appearance in the cooler stars; and a few elements disappear in passing from the hot to the cooler stars. In the hotter stars little more than hydrogen can be detected, then follow hot stars with calcium, magnesium, and a few other elements superadded; then come cooler stars with more complex spectra corresponding with a greater variety of elements. The planets, of which our own is a type, are among the cooler orbs. If the different suns and stars be arranged in a series, the order of the appearance of the elements in the cooling stars is approximately the order of their increasing complexity as deduced from the magnitude of their atomic weights. The lightest elements alone appear in the hotter stars. These facts fit very well into the hypothesis that the matter of which stars are made, passes through a real change in the nature of the constituent elements, and that there is a progressive tendency of the elements to assume more stable forms in passing from the hotter to the cooler stars. This corresponds with the assumption that the atoms are built of particles which form more and more complex aggregates as the temperature falls. Carbon is an exception. It has a low atomic weight, and yet it appears comparatively late, but the non-volatility of the solid element shows that the molecule is probably complex. Calcium (atomic weight 40) appears before sodium (atomic weight 23); this may be due to the fact that the stability of the system of corpuscles which form an atom not only depends upon the number but also upon the mode of arrangement of the corpuscles. In a general way, however, the elements appear in the cooling stars in the order of their increasing atomic weights. The stars may thus be arranged in groups corresponding with different stages in their development. The hydrogen and helium stars pass by insensible gradations into stars of the solar type, and finally into the deep red stars. Thus J. N. Lockyer (1887) states that:

	Spectrum.	Temperature.	Appearance of the elements.
Gaseous stars .	Longest	Highest	Hydrogen, helium, asterium (a gas not known on earth).
Metallic stars .	Medium	Medium	(a) Feeble spectrum of helium and hydrogen; magnesium; calcium; silicon and oxygen. (b) No gases of the helium family; iron; manganese; nickel; copper; etc.
Carbon stars .	Shortest	Lowest	Carbon and compounds of carbon.

Astronomers consider that the different stages in the evolution of sidereal systems cannot be demonstrated by the slow cooling of a *single* star, because the span of human life relative to the duration of cosmical events is far too short to enable the different stages to be followed in succession; these different stages can rather be followed by arranging the different nebulae in a series so as to show all gradations, from a diffuse luminous haze to stars with faint nebulous halos. The spectra of the gaseous and presumably younger nebulae consist of three lines corresponding with hydrogen, helium, and some unknown elements—the great

nebula in Orion is usually given as an example. As the nebulae grow older and more compact, more lines corresponding with other elements appear. These spectra are supposed to represent clusters of corpuscles more stable than the rest. Hence, according to J. N. Lockyer's evolution hypothesis, the spectra of a properly arranged series of stars and nebulae¹ indicate that the chemical atoms have grown during the cooling of the primal ultra-atomic gas much as visible rain drops grow from invisible water vapour. Before hydrogen appeared, a whole series of lighter elements were probably formed by the gradual condensation (polymerization, etc.) of cooling "fire-mist," and then passed into the heavier and more complex elements as the temperature fell still lower. From a chemical point of view, therefore, the cooling of the primal matter has resulted in the formation of a succession of polymers $(1, 2, 3, \dots)_{\eta}$ where η is the density of the hypothetical protyle. The polymerization may also proceed in successive stages. $A = (1, 2, 3, \dots)_x$; $B = (1, 2, 3, \dots)_y$; $C = (1, 2, 3, \dots)_z$; . . . where x, y, z, \dots represent the densities of simple forms of matter. These complexes unite to form the series:

$$An, Bm, Cp, \dots$$

$$AB, BC, \dots$$

$$ABC, \dots$$

where n, m, p, \dots are integers. The light elements—*asterium*, *coronium*, *nebulium*, etc.—which appear to be present in some of the more attenuated nebulae, and in the hotter stars, have probably long since vanished from the earth. We infer their existence from their characteristic spectra. Helium was once included in the list of light stellar elements unknown on earth (p. 690).

Summary.—The hypothesis which is supposed to correlate these and other facts may now be outlined. Long before the earth was formed, it is supposed that a kind of ultra-gaseous protyle was suffused throughout space, and, what has been called the "temperature" of the protyle, was inconceivably hotter than anything at present known on earth. In course of time, some process, akin to "cooling," reduced the "temperature" of the protyle so that it was condensed into material atoms. The simplest elements most nearly allied to the protyle would naturally condense first—thus, hydrogen and helium with their low atomic weights were born. Then followed the elements next in order of complexity until finally uranium or radium was born. We do not know an element with a greater atomic weight and presumably a more complex structure. If the ideas developed in the next chapter approximate to the truth, even this element is not stable, and is slowly breaking down into simpler more stable forms. As the temperature fell still lower, the earlier formed elements would unite among themselves and produce chemical compounds. It is possible to reverse the process and dissociate chemical compounds by elevating the temperature (p. 241), but it has not been possible to raise the temperature high enough to verify the hypothesis by "dissociating" the elements.

¹ There is not much room for doubt about the theory with respect to the cooling of hot stars, but with nebulae, S. Arrhenius (1907) considers that great cold reigning in space has condensed all but the lighter elements into the solid or liquid state, and these have gravitated to the interior. The outer layer only is rendered luminous by dust particles and corpuscles flying into the nebula from space.

Questions.

1. Write an account of the general chemistry of silicon, having regard to its position in Group IV.—*St. Andrews Univ.*

2. A new elementary body, a metal, is placed in your hands. What means would you take to ascertain (1) its atomic weight; and (2) its position in Mendeléeff's table?—*Science and Art Dept.*

3. Show that the chemical properties of magnesium, zinc, and cadmium justify their position in the same group in the periodic classification.—*Univ. North Wales.*

4. What was Prout's hypothesis? Write an account of the history of the doctrine to the close of the investigations of Stas, including in it an account of his views.—*Sheffield Univ.*

5. It has been asserted that in all probability every well-defined property of an element is a periodic function of the atomic weight. Explain precisely what is implied by this statement.—*Science and Art Dept.*

6. State the general characters of the magnesium-zinc-cadmium family of metals, and indicate their relations to the alkaline earths on the one hand and to aluminium on the other.—*Science and Art Dept.*

7. Reviewing the binary compounds, show that in the elements of the same family there is an increasing tendency as the atomic weight rises to form compounds having the electro-negative element in larger proportion.—*Owens Coll.*

8. What grounds had Mendeléeff for predicting the existence of the elements gallium, scandium, and germanium?—*Science and Art Dept.*

9. Do you think the following a valid criticism of the periodic law, by M. Berthelot (1885)? Referring to the preceding question, "this prediction is not a consequence of the periodic series; it results purely and simply from the laws and analogies which have been known for many years, and which are independent of the new system."

10. Compare the properties of (a) zinc and aluminium, or (b) manganese and iron, and their derivatives, and in the case of the pair selected explain why the two elements are placed in different groups in the periodic table.—*London Univ.*

11. Why is manganese included with chlorine in the same group of the periodic table? With what elements besides the halogens is manganese related, and how is this relationship displayed?—*London Univ.*

12. Describe the general characters of the family of elements of which glucinium (beryllium) is the first term in the periodic scheme.—*Board of Educ.*

13. In the periodic scheme of the elements, lead appears in the same column as tin. Justify this association by reference to the characters in which they agree.—*Board of Educ.*

14. Discuss the position of the helium family in the periodic classification.—*St. Andrews Univ.*

15. What is meant by the word "periodic" in connection with the classification of the elements? Tabulate the general characters of the family of elements usually known as the metals of the alkaline earths.—*Punjab Univ.*

16. What do you understand by the periodic arrangement of the elements? What are its chief uses? Illustrate your answer by examples. Point out any defects in this arrangement of the elements.—*Sydney Univ.*

17. In 1871, Mendeléeff predicted the existence of a metal with an atomic weight of about 69 and a density of about 6. On what basis did this prediction rest, and how was it verified?—*New Zealand Univ.*

18. What substances stand on the border-line between metals and non-metals? Illustrate the fact that such classification is always approximate, and that nearly all the laws of chemical combination or chemical relations, like Prout's law, are approximately true.—*New Zealand Univ.*

19. Give a brief historical account of the advance chemistry owes to the spectroscope. Do you know anything of attempts to correlate the atomic weights of the elements with the position of the lines of their spectra?—*Calcutta Univ.*

20. Explain how the valency of an element is connected with the position of the element in the periodic table. What are the reasons for the positions assigned to (a) hydrogen, (b) manganese or copper in the table?—*Board of Educ.*

CHAPTER XLIII

RADIOACTIVITY

§ 1. Electric Discharges in Attenuated Gases.

If a first step towards understanding the relations between æther and ponderable matter is to be made, it seems to me that the most hopeful foundation for it is knowledge derived from experiments on electricity in high vacuum.—**LORD KELVIN** (1893).

THE study of the phenomena attending the passage of electricity through gases has led to astounding developments during the past twenty years, and abundantly justified Lord Kelvin's prognostication. Under ordinary conditions, gases are such poor conductors of electricity that they are classed as good insulators. In order to get electricity to pass through air at ordinary atmospheric pressures, an electrical pressure approaching 30,000 volts per cm. is required ; and as the pressure of the air is diminished

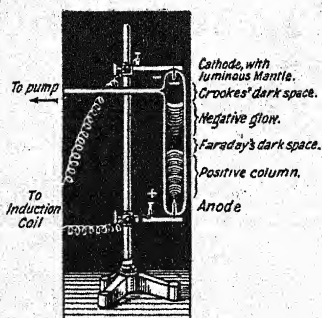


FIG. 336.

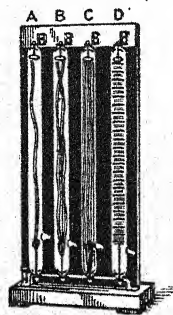


FIG. 337.

The Effect of an Electric Discharge on Attenuated Gases.

the voltage required to produce a discharge diminishes in almost exactly the same proportion. If a glass tube, about 30 cm. long, Fig. 336, be connected with a mercurial air pump, and the aluminium electrodes—disc and point—be connected with an ordinary induction coil and battery, either no spark, or thin zig-zag sparks will pass through the tube ; all depends on the distance of the electrodes apart, and on the electric pressure produced by the coil. If the pump be started, the spark passes more and more readily as exhaustion proceeds ; first, forked brush-like bluish sparks begin to leap from electrode to electrode ; when the pressure reaches 40 mm. of mercury, a luminous red streamer appears as illustrated by the tube A, Fig. 337 ; the red line widens, forming a fuzzy strip between the electrodes ;

when the exhaustion is such that the pressure is about 10 mm., a luminous band fills the whole tube—*B*, Fig. 337—and a violet halo surrounds the two electrodes. At 6 mm. pressure, the band begins to break up into layers—*C*, Fig. 337—at 3 mm. pressure, the tube appears to be filled with a number of transverse flickering reddish stræ, alternately light and dark¹; while the violet halos about the electrodes grow larger and larger—*D*, Fig. 337—and a dark space—*Faraday's dark space*—appears at the negative electrode. The prevailing colour of the pulsating stræ depends on the nature of the gas in the tube—with hydrogen the colour is red, and with chlorine green. The aureole about the cathode separates from the flickering stræ and a dark space—*Crookes' dark space*—appears between the negative glow and the cathode, Fig. 338. As exhaustion continues, the stræ diminish in number and size; and they appear to be paler in colour. The light at the anode dwindles down to a luminous point, and *Crookes' dark space* at the cathode soon expands and finally fills the tube. The glass then acquires a greenish-yellow phosphorescent light if the tube is made of soda glass. The pressure is then about 0.03 mm. of mercury. With further exhaustion the tube looks as if it were empty, but the glass still glows brightly, particularly about the cathode. With still further exhaustions, the current from the induction coil is unable to pass through the vacuum tube. The fact that the tube when highly evacuated is non-conducting shows that *the electric current must somehow be carried from one electrode to the other by something*.

§ 2. Cathode and Lenard Rays.

The electron has conquered physics, and many worship the new idol rather blindly.—*H. POINCARÉ* (1907).

Whoever rejects faith in the reality of atoms and electrons, or the electromagnetic nature of light waves, or the identity of heat and motion, cannot be found guilty of a logical or empirical contradiction; but, he will find it difficult from his standpoint to advance physical knowledge.—*M. PLANCK* (1913).

W. Hittorf (1869) showed that if a solid body—say a Maltese cross made of mica—be placed between the anode, *A*, and cathode *C*, as in Fig. 338, a true shadow appears on the glass; the shape of the cross shows that something must travel from the neighbourhood of the cathode in straight lines. This “something” which causes the phosphorescence of glass was called by *E. Goldstein* (1876) *cathode rays*. Hence, (1) *the cathode rays travel in straight lines normal to the surface of the cathode; and they will cast a well-defined shadow if a solid object be placed between the cathode and the wall of the vacuum tube*. The experiment can be

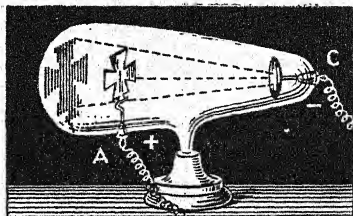


FIG. 338.—Shadows Cast by Cathode Rays.

¹ Glass tubes of about this degree of exhaustion—*Geissler's tubes*—are made in numerous patterns, and with different kinds of glass so as to get different fluorescent effects. Tubes containing gases under reduced pressure and arranged for the passage of an electric discharge are called *vacuum tubes*.

varied in an ingenious manner, as shown by W. Crookes (1879), Fig. 339, by arranging the stream of cathode rays so that it strikes the upper vanes

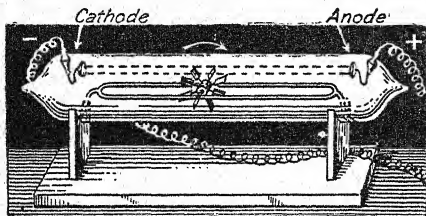


FIG. 339.—Mechanical Motion by Cathode Rays.

of a little paddle wheel which then rolls horizontally along a pair of parallel glass rails, away from the cathode. By reversing the electric current, the wheel stops and then revolves in the opposite direction owing to the fact that the direction of the cathode stream has been reversed. Hence, (2) *cathode rays can exert mechanical pressure*. By directing the cathode rays on to different minerals, beautiful phosphorescent effects may be obtained, Fig. 340. Crystals of perthite, and didymium glass give a red phosphorescence; artificial rubies, and willemite a green phosphorescence; scheelite yellow; Iceland spar, white; barium platinocyanide, zinc blende, etc., also glow and phosphoresce when exposed to the rays. Hence, (3) *many minerals become phosphorescent when exposed to the cathode rays*. The spectra of phosphorescent rare earths are of great value in studying these compounds. The cathode stream, when focused on to platinum by means of a cathode shaped like a concave mirror, may heat the metal white hot, glass can be melted, diamond charred, etc. Hence, (4) *the cathode rays raise the temperature of bodies on which they fall*. If the cathode stream be allowed to impinge on white rock salt or lithium chloride these salts are coloured violet. The glass of the vacuum tube is also coloured after

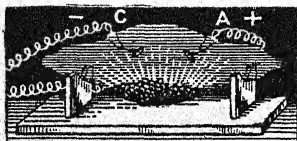


FIG. 340.—The Effect of Cathode Rays on Minerals.

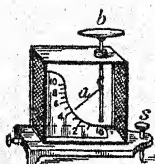


FIG. 341.—Simple Electrostatic Voltmeter.

long use. Hence, (5) *the cathode rays can produce chemical or physical changes*.

The Electrostatic Voltmeter.

The idea of the electrostatic voltmeter can be obtained from the simple form illustrated in Fig. 341. *s* is a levelling screw, gold leaf, *a*, is fixed to a rod in metallic connection with a plate *b*, all insulated from the glass box. There is a mica scale as shown in the diagram. The method of working the electrostatic voltmeter is of course described in text-books on physics. If the electrostatic voltmeter be charged with positive or with negative electrification, the gold leaf *a* will diverge from the vertical rod as shown in the diagram. If an electrified body of opposite sign be brought in contact with the plate *b*, the leaf will descend a distance proportional to the magnitude of the charge; if the electrified body is charged sufficiently to neutralize the whole of the electrification of the electrostatic voltmeter, the leaf will completely collapse; and if the electrified body has a charge in excess of that of the electrostatic voltmeter, the leaf will collapse and then ascend charged with electricity opposite in sign to that previously held. The height to which the leaf ascends is a rough measure of the magnitude of the charge. Hence, if a substance capable of charging the ambient air electrically be placed on an insulated plate in the vicinity of a charged electrostatic voltmeter, and the instrument be discharged, the rate at which the gold leaves converge will be proportional to the rate at which the air is electrically charged. Much more refined instru-

ments—electrometers—are employed in work on radioactivity where it is estimated that one-millionth of a millionth of a milligram of radium can be detected from its effect on the instrument, that is, a delicate electroscope is nearly a million times more sensitive than a spectroscope (p. 426).

In 1895 J. Perrin arranged a vacuum tube so that the cathode stream passed into a small metal cylinder inside the tube, and, by means of a wire, he connected the inner cylinder with an external electroscope. The electroscope acquired a gradually increasing negative charge, or a positively charged electroscope was discharged. Hence, (6) *the cathode rays are negatively electrified while the other contents of the tube are positively electrified.* J. Plücker (1858) showed (7) *the cathode rays can be deflected from their normal course by means of a magnet.* This is illustrated by Fig. 342, for if a magnet be directed to the side of the tube through which a discharge is passing, the focus of the rays can be deflected on to the walls of the tube. The heat produced by the bombardment of the walls of the tube by the cathode stream will suffice to melt a little wax placed on the outside of the tube, A.

P. Lenard (1894) made a tube with a thin aluminium sheet—D, Fig. 343—carried by a brass cap, at the end opposite the cathode C; a metal cylinder A served as anode. Lenard found that the cathode rays passed through the aluminium window outside the tube, and were then called Lenard rays. The cathode or Lenard rays are absorbed by different metals used as windows—the absorptive power of a substance is almost directly proportional to its density. If the metal window is too thick, the cathode stream is arrested. Hence, (8) *the cathode rays can penetrate and pass through thin sheets of metal, but not through thick sheets.*

At first, the cathode rays were thought to be a stream of negatively electrified atoms or molecules of the residual gas in the evacuated tube; C. Varley (1871), or rather W. Crookes (1879), suggested that they were particles or molecules of a fourth state of matter—an ultra-gaseous state which was called radiant matter—in which the free paths of the molecules were so long that collisions could be disregarded. Owing to their high penetrative power (Lenard rays), and the fact that no difference in the properties of electrons can be detected by changing the kind of gas in the evacuated tube, nor by changing the electrodes, it follows that (9) *the cathode rays are independent of the kind of matter present; and if the particles be matter at all, the matter is the same in kind, from whatever source it is derived.* E. Wjeczart (Jan., 1897) and J. J. Thompson

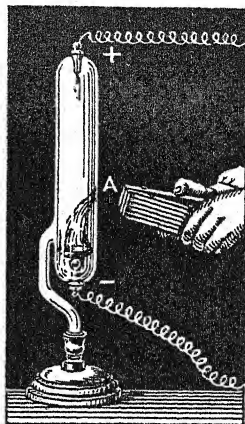


FIG. 342.—The Effect of Magnet on Cathode Rays.

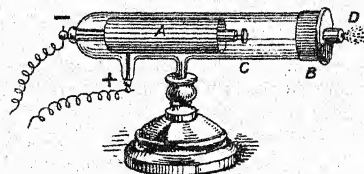


FIG. 343.—Lenard's tube to bring the Cathode Rays outside the Vacuum Tube.

(Apr., 1897) suggested the startling hypothesis that what Crookes called "radiant matter" or the cathode rays, is a stream of negatively charged particles or corpuscles which have been formed by the disintegration of atoms of the gas in the vacuum tube. The term electron was applied by G. J. Stoney (1881) to designate the unit or atomic charge of electricity, and it is now almost universally applied to the sub-atomic particles supposed to stream from the negative electrode when a discharge is passing through an attenuated gas. No difference can be detected in the corpuscles derived from different gases, and hence it is inferred that the electrons are common constituents of all gases. If a stream of electrons be directed into an atmosphere of moist air, each electron serves as a nucleus about which moisture collects, and each electron becomes the centre of a visible drop of water. Hence, (10) *a stream of electrons, when directed into moist air, forms a cloud.* The cloud or mist is an aggregate of minute falling raindrops; and it is assumed that, like a particle of dust in moist air (p. 199), an electron in moist air can serve as a nucleus for the condensation of the water vapour.

The theory of the cathode rays.—According to hypothesis, the electric discharge in attenuated gases splits the atoms of the gas into positively and negatively charged electrons. The cathode rays are a stream of negatively charged electrons sent from the cathode with a high velocity. It is inferred that ordinary atoms are probably made of nothing but aggregates of sub-atomic particles—positively and negatively charged. Under ordinary conditions, the charges counteract one another and the atom is electrically neutral. By the action of an electrical discharge, negative electrons are supposed to be detached from the atom, leaving a residue with a positive charge, and called a positive electron or positive ion. If a negative electron attaches itself to a neutral atom, the latter will acquire a negative charge. In reviewing the evidence derived from the properties of cathode rays J. J. Thompson (1897) said: "The explanation which seems to me to account for the facts in the most simple and straightforward way is founded on the view of the constitution of the chemical elements which has been favourably entertained by many chemists." The view is that the atoms of the different chemical elements contain different aggregations of particles of the same kind; otherwise expressed, that a part at least of all atoms consists of electrons.

Experiments which need not be detailed here have shown that in all probability the electric charge on an ion formed in the process of electrolysis is the same as the electric charge of an electron; that the mass of an hydrogen ion is 1700 times the mass of an electron or negatively charged corpuscle; and therefore the mass of an electron is $\frac{1}{1700}$ th the mass of a hydrogen atom—i.e. 7×10^{-28} gm., and its diameter is about 3×10^{-13} cm., whereas an atom of hydrogen has a mass of about 1.3×10^{-24} gm., and a molecule of hydrogen is about 2×10^{-24} cm., so that compared with the atom, the electron has quite a microscopic size. The electrons can travel with a velocity ranging as much as 90,000 miles per second. Their speed is dependent upon the intensity of the electrical force passing through the vacuum tube. A cathode particle travelling at this speed could go nearly twice round the earth in a second. The idea that the electrical condition of matter and its chemical activity depend upon the

addition or removal of electrons from atoms or molecules has been incorporated with the ionic hypothesis (p. 363).

§ 3. Röntgen or X-rays.

We shall never succeed in exhausting unmeasurable nature.—A. VON HUMBOLDT.

When the exhaustion of a vacuum tube is such that the tube is on the verge of becoming electrically non-conducting, and the glass opposite the cathode is brilliantly fluorescent, rays proceed from the fluorescent glass, *outside the tube*; these rays—called **X-rays** or **Röntgen rays**—have quite different properties from the cathode or Lenard rays, because they will pass through glass, and they are not deflected by a magnet. Like rays of light, Röntgen rays can be reflected, refracted, and polarized; and they are not appreciably affected by the most powerful electric or magnetic fields as charged particles would be. It is supposed that Röntgen rays—like the rays of light, radiant heat, and electro-magnetic waves—are due to pulses or waves set up in the æther by the impact of electrons on matter. It is all a question of frequency or wave-length. The following table gives the wave-lengths of a number of radiations:—

Radiation.	Wave-length in cm.
Electromagnetic waves of wireless telegraphy	3×10^3 to 5×10^4
Longest heat waves known	6×10^{-3}
Infra red spectrum	6×10^{-3} to 7.5×10^{-5}
Red spectrum	6×10^{-5}
Green spectrum	5×10^{-5}
Violet spectrum	4×10^{-5}
Ultra-violet radiations	4×10^{-5} to 2×10^{-5}
Röntgen rays	10^{-5} to 10^{-9}

Hence, the wave-length of the X-ray is about a thousand times smaller than the wave-length of sodium light, and is comparable with the size of the atom. In atmospheric air, the distance between neighbouring molecules is about 3×10^{-7} cm.

The discoverer of the X-rays, W. C. Röntgen (1895), found that they can excite fluorescence on a paper screen coated with barium platino-cyanide, BaPtCy_4 , or calcium tungstate, CaWO_4 ; they can fog a photographic plate; and make the air through which they pass a conductor of electricity. They have a remarkable power of penetrating substances opaque to ordinary light. Röntgen rays are produced by the destruction of the cathode rays and are formed when the cathode rays impinge on solid objects. Every substance when bombarded by electrons emits Röntgen rays—the glass walls of a vacuum tube, heavy metals like platinum or uranium, etc. The penetrating power of Röntgen radiations refers to the decrease in the energy of a pencil of the rays which occurs when the rays are allowed to impinge on a solid. Röntgen radiations with a low penetrative power, called *soft rays*, are emitted from a vacuum tube which has too much residual air. The supply of electrons is then plentiful;

their speed is comparatively slow; and a current of comparatively low electromotive force is needed. Conversely, radiations with a high penetrative power, called *hard rays*, are emitted if the tube be too highly exhausted. The supply of electrons is then relatively small; their speed is comparatively high; and the necessary electromotive force is high.

The rays emitted from an ordinary bulb are usually heterogeneous, mixed hard and soft, but C. G. Barkla (1906) showed that if the existing stimulus be great enough, every substance can be made to emit a set of X-rays which can be regarded as homogeneous and characteristic, in that the absorption coefficient, k , of the radiations from that substance, in some standard substance (say aluminium), is a constant, *e.g.*:

	Ca	Cr	Cu	Se	Ag	Ba
k	435.0,	136.0,	47.7,	18.9,	2.5,	0.8.

Substances with atomic weights between aluminium and silver, emit two sets of these homogeneous characteristic radiations, *e.g.* palladium emits

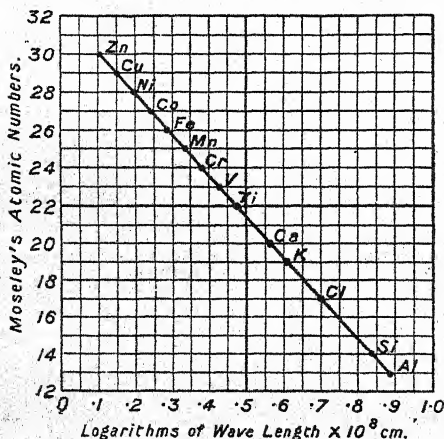


FIG. 344.—Relation between the Wave-length of the Characteristic X-ray of the Elements and the Atomic Number.

two characteristic sets of homogeneous rays with wave-lengths respectively 0.58×10^{-8} and 0.51×10^{-8} cm., and nickel, 1.66×10^{-8} and 1.50×10^{-8} cm. respectively. H. G. J. Moseley (1913) further showed that when the increase in the atomic weight of the element is plotted with the corresponding decrease in wave-length, the curve does not run smoothly; but if the logarithms of the wave-lengths or vibration frequencies be plotted against a series of natural numbers, the curve runs quite smoothly. This is illustrated by Fig. 344, in which the wave-length $\times 10^8$ cm. is

plotted against the numbers 13 to 30 ranging from aluminium to zinc for one of the homogeneous sets of rays. Analogous curves have been obtained for all the known elements. Indeed, the X-ray spectrum of every element from aluminium to gold is determined by an integer N called the atomic number ranging from 13 for aluminium to 79 for gold. There are some blanks corresponding with unknown elements. The order of the atomic numbers is the same as the order of the atomic weights, except where the latter disagrees with the order of the chemical properties so that the atomic number in the periodic table is a more fundamental index of quality than the atomic weight. This shows that the wave-length, or the vibration frequency, of the characteristic X-rays from different elements changes from element to element by regular jumps. The steady decrease in the wave-length of the characteristic X-rays of a series of elements in the periodic table depends on some fundamental property of the atoms. As

an hypothesis, it has been suggested that the increments are due to the successive addition of a unit electric charge to the nucleus of the atom; and that the magnitude of the nucleus of an atom is proportional to a number indicating the place of the element in the periodic table—hydrogen has a nucleus charge of one unit, helium two, etc.

The atomic weights of the elements do not increase in an orderly way, they mount by steps of two, but not very regularly, and the elements sometimes appear to get in the wrong order, *e.g.* nickel comes *before* cobalt, but certain chemical properties and the wave-length of the characteristic X-rays from nickel show that it occupies a position *after* cobalt. When the elements are arranged in this way the series shows gaps between cerium and gold, presumably waiting to be filled by elements yet undiscovered.

§ 4. Positive or Canal Rays.

The corpuscular theory of the positive rays has no other justification than that it explains our present-day knowledge of the phenomenon in the simplest possible way.—E. GEHRCKE.

When a perforated cathode is employed in the vacuum tube for producing cathode rays, E. Goldstein (1866) first noticed that streams of

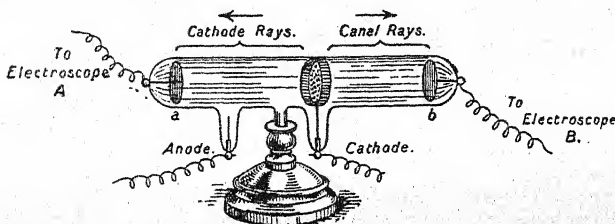


FIG. 345.—Tube for Canal or Positive rays. (After W. Wien.)

violet light passed through the perforations or canals and emerged behind the cathode on the side remote from the anode, and hence he called these streams *Kanalstrahlen* (canal rays). In the apparatus illustrated in Fig. 345, the canal rays travelling towards the right strike against a plate *b* connected with an electroscope *B*, and there show a positive charge; the cathode rays travelling towards the left, impinge on the plate *a* connected with an electroscope *A*, and there show a negative charge. The canal rays have been investigated by methods similar to those employed for the cathode rays. The results indicate that the canal rays are streams of particles the majority of which are positively electrified—hence, the term *positive rays* is replacing the older term *canal rays*. The streaming particles travel in straight lines and produce a phosphorescence (usually violet) when they impinge upon glass, etc. The speed of the positive electrons is usually much less than that of the negative electrons; and they are not so sensitive to magnetic influences. Measurements similar to those employed for the cathode rays show that the positively charged particles must be of atomic dimensions, and in no case is the mass of the positive ion perceptibly less than that of the hydrogen atom. Remembering that, so far as we can tell, all electrons are the same, and have a mass 1700 times

less than that of the hydrogen atom, while the mass of the positively charged particles depends upon the nature of the gas and is virtually the same as that of the atom from which it is derived, it is probable that when a gas is ionized, one or more negatively electrified particles—electrons—are expelled from the atom, and the corresponding positively charged nucleus remains.

J. J. Thomson's positive ray analysis.—If a mixture of different kinds of electrified particles travelling at a high velocity in one thin stream be passed simultaneously through electric and magnetic fields disposed at right angles to one another, the different kinds of particles are sorted into a number of smaller separate streams. The various streams can be demonstrated by causing them to impinge upon a fluorescent screen of, say, willemite, or on suitable photographic plates. J. J. Thomson calls the series of small streams obtained from the original stream, the electric spectrum of the atoms, for the stream of electrified atoms is separated by this treatment into sub-streams—much as a beam of light from an illuminating gas is analyzed into different rays by passing through the prism of a spectroscope. From the displacement of the path of the stream of electrified particles under the influence of electric and magnetic forces of known intensity, J. J. Thomson has developed what he calls positive ray analysis; this furnishes evidence on such subjects as:

Are the atoms of an element all alike?—The particles in any one stream are presumably all of the same kind because the parabolic curves are sharply divided, and show no tendency to merge one into another. If it were otherwise the curves would be "fuzzy." Ordinary chemical analysis cannot decide whether particles of one kind of substance have all the same mass, because it can deal only with the average masses of billions of particles.

Atoms and molecules with multiple charges.—Any given gas furnishes a number of different curves showing the presence of positive ions with multiple charges. Neutral molecules and atoms give spots, not curves. Thus, with oxygen, in addition to (1) ordinary neutral molecules, O_2 ; and (2) neutral atoms, O ; effects were obtained corresponding with streams (3) of atoms with one positive charge, O^+ ; (4) of atoms with two positive charges, O^{++} ; (5) of atoms with one negative charge O^- ; (6) of molecules with one positive charge, O_2^+ ; (7) ozone with one positive charge O_3^+ ; and (8) molecules O_6 with a positive charge, O_6^+ . With hydrogen, evidence of entities H , H_2 , H^+ , H^- , and H_2^+ was obtained.

Intermediate stages of chemical action.—The photographic plate registers the rays within a millionth of a second after their formation, so that if a chemical reaction were taking place in the tube, it is possible that the method would disclose the existence of transient intermediate compounds as well as the final products of the reaction. For example, with methane, CH_4 , five lines occur corresponding with particles having masses 12, 13, 14, 15, and 16. These must correspond with particles having the composition, C , CH , CH_2 , CH_3 , and CH_4 . Phosgene, $COCl_2$, furnishes lines corresponding with particles having masses 99, 28 and 35.5, hence the decomposition proceeds by a separation of chlorine atoms from carbon monoxide without rupture of the carbon and oxygen atoms. No signs of a molecule NO_2 were observed during the oxidation of nitric oxide, NO , by oxygen.

Atomic weights.—The parabolic tracks recorded on the photographic plates enable the atomic weight of a gaseous substance to be determined within one per cent. of its true value without requiring more than 0.00001 gram of the substance. The result, moreover, is not dependent upon the purity of the material, for impurities merely produce additional lines in the positive ray spectrum, and do not affect the curves produced by the substance under examination.

New elements.—If a spectroscopist observed an unknown line in the spectrum of a discharge tube, he would infer the existence of an unknown substance provided the line were not produced by some alteration in the condition of the discharge; similarly, if a new curve be obtained in the positive ray spectrum, the probable existence of a new element would be inferred. Thus, atmospheric nitrogen gives a curve corresponding with a substance having an atomic weight 40 times that of a hydrogen atom, and is not indicated on the plate when chemically prepared nitrogen is employed. The positive ray method, too, is far more delicate than spectrum analysis, for it enables a foreign gas to be detected in quantities too minute to be revealed by the spectroscope. Thomson found unknown lines indicating that the gas which has been called neon is probably a mixture of two different gases of atomic weights 20 and 22—but with a preponderating amount of the former. Many of the ordinary elements appear to be composed of two or more components which cannot be separated by ordinary chemical processes, and F. Soddy called them isotopic elements or isotopes (*vide infra*). F. W. Aston has greatly improved the original positive ray analyzer, and applied the principle to a number of elements; as a result, he found that *hydrogen, helium, nitrogen, oxygen, carbon, sodium, arsenic, phosphorus, scandium, titanium, vanadium, chromium, manganese, cobalt, strontium, yttrium, and fluorine* do not show isotopic forms; *boron* has two isotopes 10, 11; *lithium*, 6, 7; *potassium*, 39, 41; *caesium*, 85, 87; *neon*, 20, 22; *argon*, 36, 40; *sulphur*, 32; *chlorine*, 35, 37; *bromine*, 79, 81; *copper*, 63, 65; *gallium*, 69, 71; and *silver* 107, 109; *magnesium* has three isotopes, 24, 25, 26; *germanium*, 70, 72, 74; *silicon*, 28, 29, 30; *krypton*, 78, 80, 82, 83, 84, 86; *xenon* and *mercury* have many isotopes not yet determined exactly.

Allotropic hydrogen.—Thomson finds that the gas with atomic weight 3 is given off by most solids when bombarded by the cathode rays. It is interesting to remember that D. I. Mendeléeff (1871) predicted a new element of the halogen group with an atomic weight 3, but Thomson thinks that the gas in question is a triatomic molecule of hydrogen, H_3 , mainly because (1) deliquescent salts or salts containing combined hydrogen—*e.g.* KOH , $CaCl_2$, $LiOH$ —give continuous yields of the gas, while the supply with salts which do not contain combined hydrogen—*e.g.* LiI , Li_2CO_3 , KCl —is soon exhausted. (2) Attempts to obtain spectroscope evidence of the new gas gave bright hydrogen lines with traces of mercury—derived from the apparatus used in manipulating the gas. (3) Vigorous sparking in the presence of oxygen, or contact with glowing copper oxide (or even exposure to bright light) destroys the gas. Assuming the gas is really H_3 , it is more stable than ozone, more stable indeed than any known allotropic form of an element. If hydrogen is univalent, it is difficult to reconcile its existence with the ordinary views about valency. Thomson explains it by assuming that the hydrogen atom with its positive nucleus and negative corpuscles

exerts a force analogous with that exerted by a magnet ; and a group of three atoms can arrange themselves about their axes to form a closed stable ring.

§ 5. Becquerel Rays.

The secret of all who make discoveries is to look upon nothing as impossible.
—J. VON LIEBIG.

About the time Röntgen (1895) discovered the peculiar X-rays radiating from phosphorescing Crookes' tubes, H. Becquerel (1896) repeated some experiments of Niepce de St. Victor (1867) in order to find "if the property of emitting very penetrative rays is intimately connected with phosphorescence." In other words, does the principle of reversibility (pp. 38 and 241) apply? If Röntgen rays make a fluorescent substance shine in the dark, will a fluorescent substance emit invisible penetrative rays? Becquerel placed fragments of several phosphorescent substances on photographic plates wrapped in two sheets of black paper. In about 24 hours, when the plates were developed, a silhouette of the phosphorescent substance appeared on the plate. Hence, it was inferred that "*the phosphorescent salts of uranium must emit radiations which are capable of passing through black paper opaque to ordinary light, and of reducing the silver salts of the photographic plate, even when the uranium compound has been completely sheltered from the light.*" The radiations emitted by the phosphorescent substance are called Becquerel rays, though Niepce, thirty years previously, noticed that uranium salts could effect photographic plates in the dark, and G. le Bon (1896) called the radiations *lumière noire*—"black light." A substance which possesses the property of emitting these penetrative rays is said to be "radioactive," and the property itself is called **radioactivity**. All substances containing uranium are radioactive. The energy is not stored like light in a phosphorescent substance, because the property is no different whether insulated or non-insulated uranium be used; the energy does not come from the air because it is not affected by confining the uranium in a vacuum. It is therefore probable that the uranium is slowly undergoing some spontaneous change as an effect of its internal energy. Radioactivity does not depend upon light or heat; the emission of the rays appears to be a permanent and abiding property of uranium and its compounds; and it is independent of temperature and of all known physical conditions. No sign of a diminution or increase of the property has been detected whether the substance be heated towards 2000° or cooled towards -200°. The same weight of uranium, no matter how combined, emits the same amount of radiation. *The chemical properties of the elements—excepting perhaps the helium family—can be modified and controlled by changes in the chemical and physical conditions; but radioactivity is independent of these conditions.* Consequently, the astounding assumption is made that radioactivity is an **infra-atomic** property, and is not the same type of phenomenon as an ordinary chemical reaction. If chemistry be confined to the study of phenomena with the atom as unit, radioactivity regarded as an infra-atomic phenomenon, is a kind of *meta-chemistry*.

Becquerel also found that when uranium is brought near to a charged gold-leaf electroscope, Fig. 341, the gold leaf gradually collapses. The rate at which an electroscope is discharged is a measure of the efficiency

of the specimen in emitting rays. The charged electroscope, indeed, is more sensitive than the photographic plate for detecting Becquerel rays. Air which has been in contact with uranium and its compounds, like air which has been exposed to Röntgen rays, will discharge an electroscope, for exposure to these radiations makes air a conductor.

G. C. Schmidt (1898) found that thorium is radioactive in the same sense that uranium is radioactive, and curiously enough, these two elements have the highest atomic weight—Th, 232; U, 238. The radioactivity of thorium is readily shown by flattening an ordinary new gas mantle on the sensitive side of a photographic plate, and leaving all in darkness for about a week. When the plate is developed in the usual way, a photograph of the flattened mantle will be produced. The mantle contains sufficient thorium, as oxide, to demonstrate the effect.

It must be added that there are several strictly chemical reactions—hydration of quinine sulphate; action of sodium amalgam or calcium carbide on water; oxidation of phosphorus; combustion generally; etc.—which have the power of rendering the ambient air a conductor of electricity so that it can discharge a charged electroscope; and an attempt has been made to find whether the two phenomena are related. The temperature at which hydrogen and oxygen begin to combine in contact with carbon or platinum is almost the same, within the limits of the errors of measurement, as that at which these elements begin to form negatively charged electrons. This, however, is not sufficient to justify a belief that there is a *causal nexus* between the two phenomena. The property exhibited by many chemical reactions of making the ambient air electrically conducting must be sharply distinguished from radioactivity. Rutherford has emphasized the fact that the activity of radioactive bodies has these special characteristics: (1) It is spontaneous; (2) It is exhibited by all the compounds of the radioactive elements; (3) It is not altered by changes in the physical or chemical condition of the element. It might also be added that exposing metals to ultra-violet light, heating metal wires, splashing liquids, etc., also makes the ambient air electrically conducting.

§ 6. Radium.

It is the glory of God to conceal a thing, but the glory of a king to search it out.—SOLOMON.

For even the things which be in our hands—
These, knowing, we know not—so far from us,
In doubtful dimness, gleams the star of truth.

ANON.

Is the radioactivity of uranium due to the presence of an impurity? —P. and S. Curie tried to answer this question by examining the radioactivity of a number of uranium minerals. They found the following relative results:

	Units		Units
Pitchblende (Jochimstahl)	7.0	Uranium oxide (green)	1.8
Carnotite	6.2	Pitchblende (Cornish)	1.6
Chalcolite	5.2	Thorium oxide	0.1 to 1.4
Metallic uranium	2.3	Uranium nitrate	0.7
Orangite	2.0	Monazite	0.5

Obviously some uraniferous minerals are more active than uranium

itself. Hence, it was inferred that "the strong activity of the pitchblende from Joachimstahl (Bohemia) is due to the presence of small quantities of a substance wonderfully radioactive, and different from uranium or any other simple body known." This result was confirmed by the extraction of the chloride of what was supposed to be a new element designated radium, Ra. The salt was over a million times more radioactive than uranium. One ton of pitchblende contains about 0.37 gram of radium, and about half of this is obtained by the following method of extraction.

The extraction of radium bromide or radium chloride from pitchblende.—The pitchblende is roasted with sodium carbonate and digested with dilute sulphuric acid; the solution contains uranium, and the residue contains radium and impurities. The residue is boiled with sodium hydroxide; washed with water, and then with dilute hydrochloric acid. The insoluble residue is digested with sodium carbonate to convert the sulphates to carbonates. Wash the insoluble carbonates with water, and dissolve the mass in hydrochloric acid. The solution contains radium, etc. Add sulphuric acid to precipitate the radium, etc., as sulphates. Again digest the precipitate with sodium carbonate, wash with water, and dissolve in hydrochloric acid. Hydrogen sulphide will precipitate radioactive polonium. The ton of pitchblende furnishes about 0.00004 gm. of this element. Oxidize the solution with chlorine and add ammonia, radioactive actinium is precipitated. The solution is boiled with sodium carbonate, washed with water, and evaporated to dryness with hydrobromic acid. On the addition of hydrobromic acid, radium and barium bromides are precipitated. These are separated by fractional crystallization.

The properties of radium and its salts.—The chemical reactions of radium chloride are so like barium chloride that a separation is very difficult, and a slight difference in solubility is the only means of separating the two. The bromides are more easily separated than the chlorides. The spectrum is characteristic and related to that of the alkaline earths. Metallic radium has been isolated by the electrolysis of radium chloride with an anode of iridiumized platinum and a mercury cathode. The radium amalgam so obtained was heated in a current of hydrogen to volatilize the mercury. A white metal remained. This melted at about 700°. The metal turns black in air, possibly owing to the formation of a nitride; it also chars paper; dissolves rapidly and completely in water and in dilute hydrochloric acid, thus showing that the oxide is soluble in the solvents named. 0.0919 gram of the anhydrous chloride, whose spectrum showed the presence of barium but faintly, gave 0.0859 gram of silver chloride. Hence, 0.0919 gram of radium chloride contains the equivalent of 0.0213 gram of chlorine; or 35.46 grams of chlorine unite with 117.5 grams of radium. Hence, the equivalent of radium is 117.5. If radium chloride be RaCl_2 , by analogy with barium chloride, BaCl_2 , the atomic weight of radium is nearly 235. Later determinations give rather lower values—226.5.

Radium appears to be a member of the family of alkaline earths. Its salts resemble the corresponding salts of barium. Radium sulphate is less soluble than the barium salt, the carbonate also is sparingly soluble. The chloride and bromide crystallize with two molecules of water: $\text{RaCl}_2 \cdot 2\text{H}_2\text{O}$; $\text{RaBr}_2 \cdot 2\text{H}_2\text{O}$, and these crystals are isomorphous with the corre-

sponding barium salts. The radium halides are much less soluble than the barium halides, and this enables radium to be separated from barium. Several other salts—nitrate, azoimide, cyanoplatinate, etc.—have been prepared. The radium salts when freshly prepared are white, but they afterwards become yellow and brown, particularly if the salts are impure. Solutions of radium salts have a blue luminescence, and the salts are all luminous in the dark. The spectrum of radium is characteristic, and it exhibits lines which belong to no other known element. Radium salts impart a crimson coloration to Bunsen's flame.

The chemical effects of the radiations from radium.—The Becquerel rays evolved from radium chloride closely resemble those from uranium and they produce similar effects, but over a million times more intense. The Becquerel rays from radium chloride or radium bromide incite phosphorescence in diamonds, rubies, fluorspar, calcium sulphide, zinc sulphide, barium platinocyanide, etc. If the eyes be closed, and a tube of radium bromide be held near the forehead, the retina of the eye becomes phosphorescent, and light will be seen though the eyes are closed. A tube containing a little radium bromide when held near the skin for a few hours produces painful sores. Caterpillars and other small animals are said to be killed if shut up in a box with a minute fragment of radium. The radiations coagulate proteid matter—*e.g.* globulin. It is also claimed that the exposure of malignant skin diseases, superficial cancer nodules, etc., has proved beneficial in many cases, although the testimony of medical experts is not unanimous. Becquerel rays cause chemical action—discolour paper and glass; turn oxygen into ozone; form hydrogen peroxide in acid, neutral, or alkaline solutions; decompose water; convert yellow phosphorus into the red variety; reduce mercuric to mercurous chloride, and ferric to ferrous sulphate; decompose iodoform, potassium bromide, hydrogen sulphide, carbon dioxide, etc. An aqueous solution of a radium salt continuously evolves hydrogen and oxygen gases, owing to the decomposition of the water. The radiations affect photographic paper, and discharge an electroscope as already indicated.

The action of a magnet on the radiations from radium.—A few sheets of paper or a couple of sheets of aluminium foil will cut off a large part of the radiations, and a sheet of lead, about half a centimetre thick, will cut off nearly all the radiations. A residuum still remains unsuppressed even after passing through 15 cm. of lead or through a far thicker block of iron. Hence, the radiations from radium are not homogeneous. Again, the radiations from radium are not affected in the same way by a magnet. Some of the rays are not influenced, for they do not bend when placed in a magnetic field, these are called the γ -rays. Others are bent *towards* the magnet, and are called the β -rays; while others are bent *away* from the magnet, and are called the α -rays. The three distinct types of rays in the radiations emitted from radium salts are as follows:

1. Alpha rays.—The α -rays are slightly bent by intense magnetic forces; they have a positive charge; and slight penetrative power, so that they are suppressed by a few layers of paper or a few cm. of air. The general properties of the α -rays correspond with those of the canal rays in a vacuum tube. The experimental evidence leads to the inference that the α -rays are streams of positively charged electrons projected

from radium with a velocity approaching 20,000 miles per second. The emission of α -particles from radium salts can be illustrated very neatly by W. Crookes' spinthariscopes. A small fragment of a radium salt supported at the tip of a wire, *B* (Fig. 346), in front of a screen, *A*, coated on the inside with zinc blende, is viewed in the dark through a magnifying eyepiece, *E*, which is focused on to the screen by sliding it up or down the tube. Flashes of light are continually scintillating on the screen. R. K. Duncan likens the effect to the appearance of a swarm of fireflies on a dark night. The scintillations are caused by the rain of α -particles from the radium salt on to the screen. Each impact is marked by a flash of light. Each α -particle is supposed to produce one flash.

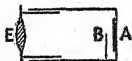


Fig. 346.—Spinthariscopes.

2. Beta rays.—The β -rays are readily bent by comparatively weak magnetic forces in the opposite direction to the α -rays. The β -rays have a negative charge; and a stronger penetrative power than the α -rays. The general properties of the β -rays correspond with those of the cathode rays of the vacuum tube, for they appear to be negatively charged electrons or corpuscles projected from the radium salt with a velocity approaching 100,000 miles per second. The corpuscles in Crookes' tube travel a little more slowly than the β -rays from radium. A difference in speed might be expected from their different modes of generation. The

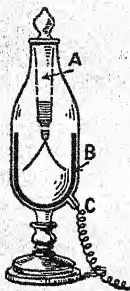


Fig. 347.—Strutt's Radium Clock.

following experiment—R. J. Strutt's radium clock—illustrates the character of the β -rays. A morsel of a radium salt is enclosed in a glass tube, *A*, Fig. 347, coated with a conducting material and ending at the bottom with a brass cap from which hang a pair of gold leaves. This system is fitted within a glass tube from which the air is exhausted. The inside of the tube is coated with strips of tinfoil, *B*, connected with the earth by the wire *C*. The β -rays from the radium pass through the glass and leave the central system with a positive charge. This causes the gold leaves to gradually diverge until they touch the tinfoil, when they are discharged, and the leaves collapse. The process begins anew. This charge and discharge goes on continuously since the radium can emit these radiations an indefinitely long time. This arrangement is perhaps the nearest approach yet made to perpetual motion. The frequency of the cycle, of course, depends upon the amount of radium in the inner tube.

3. Gamma rays.—The γ -rays are not affected by the most intense magnetic forces. Their penetrative power is very intense, and they can manifest their presence after passing through several inches of metallic lead or several feet of metallic iron. The relative penetrative powers of the three types of rays for aluminium are roughly as $\alpha : \beta : \gamma = 10 : 10^3 : 10^6$. The γ -rays do not appear to be material particles at all, but the experimental evidence shows that the γ -rays are similar to, if not identical with, Röntgen rays. A diagrammatic illustration of the three types of radiation from radium can now be given, Fig. 348. A piece of radium is supposed to be placed in a lead vessel, *A*, sufficiently thick to prevent rays travelling through the walls. Under the influence of an intense magnetic

field, the rays no longer travel in straight lines, but they are deflected as shown diagrammatically in the figure.

The spontaneous degradation of energy by radium.—All three types of rays are continuously emitted by radium compounds in their normal condition. The intense radioactivity of radium appears to be associated with the fact that the temperature of the salts is always a little—about 1.5° —above the temperature of their surroundings. Normally, one gram of radium appears to evolve enough energy to raise a little more than its own weight of water from freezing to boiling point every hour; this amount of heat is equivalent to 118 Cals. per hour, or about 1000 cals. per annum. It has been estimated that a gram of radium will continue radioactive for about 2500 years, and it therefore follows that a gram of radium gives energy equivalent to that obtained by burning $\frac{1}{10}$ ths of a ton of coal during the period of radioactive change. Otherwise expressed, radium furnishes 250,000 times as much energy as is given by burning an equal weight of coal. Obviously, too, a relatively large amount of energy is needed for the continuous decomposition of water by radium salts in aqueous solutions. Hence radium is continually doing work at an undiminishing speed without any external supply of energy; otherwise expressed, the reaction is exothermal. Whence comes this supply of energy?

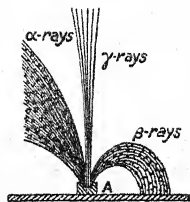


FIG. 348.—Diagrammatic Analysis of the Radium Radiations by a Magnet.

§ 7. The Degradation of Radium.

All things have their emanations.—EMPEDOCLES.

Matter, formerly regarded as inert and only able to give back energy originally supplied to it, is, on the other hand, a colossal reservoir of energy—infra-atomic energy—which it can expend without borrowing anything from without.—G. LE BON (1908).

How does the continuous evolution of heat by radium agree with the dogma that heat cannot come from nothing, but must come from some other source? The heat of a stove is derived from the oxidation of the fuel inside; with radium, it is assumed that the atoms (or molecules) are continually changing. If radium be an element, and the radium atoms are changing into something not radium, it follows that there must be some flaw in the hitherto universally accepted definition of an element.

According to P. and S. Curie:

Any substance placed in the neighbourhood of radium acquires a radioactivity which persists for many hours and even days after the removal of the radium. This induced radioactivity increases with the time during which it is exposed to the action of radium up to a certain limit. After the radium is removed, it decreases rapidly and tends to disappear. The kind of substance exposed to the radium is almost a matter of indifference, for all substances acquire a radioactivity of their own.

This fact has been traced to the continuous evolution of a substance from radium which behaves as if it were a radioactive gas. The emanation emits only the α -rays, that is, positive ions virtually as large as atoms

themselves. To avoid the hypothesis implied by calling the radioactive substance, a "gas," E. Rutherford, its discoverer, called it an *emanation*.

The radium emanation is quite distinct from the three types of rays emitted by radium and its salts. The amount is very small, but the supply is continuous. If the temperature be raised the absorbed emanation is given off, for a short time, much more copiously than in the cold. There is now little doubt that it is a gas, for it has a characteristic spectrum, somewhat resembling the spectrum of xenon; it can be condensed by liquid air to a minute drop of liquid (microscopic) of specific gravity 5.7; and at still lower temperatures, the liquid solidifies. The solid melts at about -71° , and the liquid boils at about -62° . The radioactive gas is chemically inert, for it resists attack by every chemical reagent hitherto tried, it has a characteristic spectrum, and distinctive chemical and physical properties, and in consequence, the radium emanation has been placed with the argon-

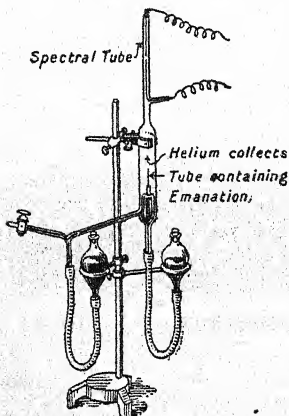


FIG. 349.—Rutherford's Experiment showing the Change of α -particles into Helium.

helium group of the periodic table. The atomic weight of the emanation, if it has an atomic weight, appears to be nearly 222.5 ($H_2 = 2$). This has been determined by the gas density, the diffusion, and the effusion methods. It has been suggested that the gaseous radium emanation be called niton, Nt, from the Latin *nitens*, shining.

The emanation, if kept by itself, slowly disappears. After about four days, only about half the original quantity remains. In fact, the radium emanation decomposes, continuously and spontaneously, into a radioactive solid and helium gas. This latter observation is important. Rutherford enclosed some radium emanation in a thin-walled glass tube surrounded by a vacuum jacket, Fig. 349. Each vessel was gastight; α -rays from the radium emanation could penetrate the walls of the inner vessel, but not the walls of the outer tube. By raising the left mercury reservoir, the gas in the annular space could be compressed in the spectrum tube, and there sparked, and examined spectroscopically. In two to six days' time, a gas with the spectrum of helium accumulated in the annular space between the two tubes. To show that the helium was not derived by diffusion from the inner tube, the emanation was removed and helium substituted. No trace of helium could be detected in the outer vessel after standing several days. This and other experiments have proved that the radium passes into helium *via* the emanation. Otherwise expressed, helium is one decomposition product of radium. Still further, the α -rays are streams of positively charged particles, each particle carries two unit positive charges, and each α -particle is an atom of helium carrying two unit positive charges.

The action of the radiations on air.—The movement of the swiftly speeding α -particles is obstructed as they pass through a gas; and the kinetic energy (p. 135) is expended in doing work on the gas. The gas

into which the α -particles are discharged becomes a conductor of electricity, and it acquires a greater chemical activity. The gas is said to be ionized. The α -particles can travel but a few centimetres before their initial velocity is exhausted, and they pass into ordinary molecules of helium. In air, for instance, the α -particles from uranium are slowest—they can travel about 2.5 cm.—whereas the α -particles from thorium C_2 travel fastest—about 8.6 cm.

The career of the α -particles has been investigated at different stages of its flight. At the beginning, when the speed and kinetic energy are greatest, the particle does *least* work in ionization. It is supposed that just as a swiftly speeding bullet can cut a clean hole in a pane of glass, while a slowly moving bullet will shatter the glass, so when the speed of the α -particles is greatest, they can pass right through the molecules of a gas without producing any appreciable effect; but towards the end of their flight the ionizing power is greater because of the greater shattering power of the slowly moving molecules; and finally the particle ceases to ionize the gas and settles down to "ordinary life" as a helium molecule (or atom). The stopping power of a gaseous medium is determined, not by the physical or chemical properties of the gas, but solely by its molecular weight.

The path of the ray through a moist gas has been made visible, owing to the condensation of water about the ions produced in the track of each α -particle. C. T. R. Wilson (1912) has photographed the fog so produced in the tracks of the α -particles, and the photographs make it very plain that the ionizing power of the particles rapidly reaches a maximum, and then rapidly sinks to zero as the molecules of helium. Similar fogs produced by the β -rays have been photographed. If a β -ray moves rapidly it produces a straight track, but the slower-moving rays have more or less tortuous tracks, being deflected hither and thither by collision with the molecules of the gas. The tracks produced by the Röntgen rays closely resemble those produced by the β -rays, and it is presumed that the Röntgen or γ -rays excite β -rays in matter on which they fall, and that the effects produced by the γ -rays are directly due to the β -rays to which they give birth.

Counting the α -particles emitted by radium.—E. Rutherford and H. Geiger (1908) devised an electrical method for counting the α -particles.

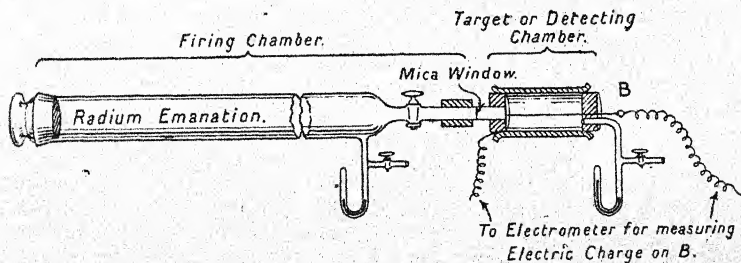


FIG. 350.—Apparatus for counting the α -particles emitted by radium emanation. (After E. Rutherford and H. Geiger.)

A small disc with a little emanation was fixed to a small iron cylinder and placed in a long glass cylinder—Fig. 350. The position of the

radium disc could be adjusted by means of a magnet. This chamber—called the *firing chamber*—had a small mica window opening into a brass cylinder with ebonite ends and fitted with a central wire, *B*, connected with an electrometer. The brass cylinder was exhausted, and it then forms the target or *detecting chamber*. It was found possible to adjust the conditions so that two to five particles per minute passed through the mica window and upset the electrical equilibrium of the testing chamber. Each particle signalled its arrival by an unmistakable jump of the needle of the electrometer. The needle can be arranged so that its movements are recorded on a chart (Fig. 351). Each jump represents the arrival of one particle in the “detecting chamber.” Some of the jumps show that two particles sometimes arrive simultaneously or almost simultaneously. Given the size of the window and the distance of the source of the projectiles, simple arithmetic furnishes the number shot per minute in all directions from the radium emanation. The results were consistent with those obtained by counting the number of scintillations produced per minute when a zinc sulphide screen was used as a target, and show that about 3.4×10^{10} of α -particles are shot per second from a gram of radium. Each α -particle represents one atom of helium. Further experiments also



Fig. 351.—Chart showing movements of the needle of the electrometer on the arrival of α -particles.

showed that a gram of radium furnishes about 0.158 c.c. of helium per year, hence the weight of an atom of helium can be computed.

The restoration and decay of the radioactivity of radium.—The radium from which the emanation has been abstracted loses about 75 per cent. of its radioactivity and it then emits practically nothing but the α -rays; the β - and the γ -ray activities are almost completely lost. The normal radioactivity of radium gradually returns to its original value on standing. The rate at which the exhausted radium regains its activity is equal to the rate at which the emanation loses its radioactivity. Hence it is inferred that radium is constantly generating and storing the emanation, and that the emanation is constantly decaying. We have here the principle of opposing reactions, and the radioactivity of normal radium is an equilibrium value because the rates of production and disintegration of the emanation are evenly balanced. The processes of decay and restoration cannot be influenced by any known controllable physical or chemical force¹; they are independent of the chemical form of radium—chloride, bromide, carbonate, sulphate, metal, etc.; all we can do is to study the mode and measure the rate of change. Hence, rightly or wrongly, it has been inferred that the process is a property of the radium atom alone; that the radium atoms break down into atoms of the emanation; and the atoms of the emanation break down into a radioactive solid

¹ This has, of course, no reference to the statement in the text to the effect that the emanation can be driven off from radium with a greater velocity at elevated temperatures because the emanation is already there. The rate at which the radium “manufactures” the emanation is not affected by temperature, etc.

and helium gas. This reminds us of the step by step decomposition of potassium hypochlorite into potassium chlorate, perchlorate, and chloride as well as oxygen. By analogy with the evolution of heat which attends certain exothermal chemical reactions—*e.g.* the decomposition of ozone $2O_3 \rightarrow 3O_2 + 68.2 \text{ Cals.}$ —it has been stated as an hypothesis that the origin of the energy of radioactive bodies is due to the decomposition of the atoms into electrons. In the one case, the atoms liberated during the reaction recombine to form new and more stable molecules; and in the other case, the liberated electrons unite to form new and more stable atoms. According to this hypothesis, a radioactive substance must have a more or less limited period of existence or life.

The products of the decay of radium.—The solid radioactive deposit has been studied, and it has been found to “decay” into a series of products some of which give α -, β -, or γ -rays, and the presence of the radioactive solid is therefore the source of the β - and the γ -rays of radium. Hence, radium normally contains all three products—radium proper, the emanating gas, and the radioactive solid. It is estimated that 25 per cent. of the radiations of normal radium belong to radium proper; 18 per cent. more of the α -rays belong to the emanating gas; and the remaining 57 per cent. to the radioactive solid. The products of the decay of radium—radium-A; radium-B; etc.—so far as they have been at present made out, are as follows:

THE DESCENDANTS OF RADIIUM.

Radium	A	→	B	→	C ₁	$\begin{matrix} \nearrow D_1 (1.4 \text{ min.}; \\ \beta\text{- and } \gamma\text{-}) \\ \searrow D_2 \end{matrix} \rightleftharpoons \begin{matrix} E_1 \\ E_2 \end{matrix}$	→	F
Average life ¹ . . .	3 m.		26.8 m.		19.5 m.	15 y. ?	4.8 d.	140 d.
Radiation	α-		β-		α-, β-, γ-	β-	β-, γ-	α-rays
Estimated atomic weight . . .	218		214		214	210	210	210

Helium is undoubtedly formed at each stage of the degradation. Radium-F appears to be the same as S. Curie's radioactive polonium which, in the absence of proof to the contrary, is thought, on further change, to pass into lead.

F. Soddy and H. Hyman (1914) tried whether the atomic weight of lead extracted from the mineral thorite and ordinary lead showed any difference. They reported that the former gave the smaller result, though no special precautions were taken to eliminate silver. M. Curie (1914) similarly found that lead derived from pitchblende had an atomic weight 206.5 when galenic lead had the value 207.01. Hence, it has been inferred that “there are several varieties of lead of different atomic weights determined by the initial metal from which they are derived.” In general, lead derived from radioactive minerals has a lower atomic weight than ordinary galenic lead. T. W. Richards and R. P. Calvert (1914) add:

The inference seems to be that radioactive lead contains an admixture of some substance different from ordinary lead, and very difficult to separate from it by

¹ The term “average life” employed in the study of radioactivity corresponds with the following illustration: “If a church at a fixed time contains a number of people of different ages, then, given the number of years each person will live after leaving the church, the average life of the congregation is reckoned from the time fixed by that church attendance.” This is not the average life taken from the time of birth used in computing insurance rates.—A. T. CAMERON (1910).

chemical means. This substance cannot be identified in the ultra-violet spectrum of the material, either because it has the same spectrum as lead, or because it has no spectrum in that part of the field, or because its spectrum is masked or absorbed by that of lead.

With the elements copper, silver, iron, sodium, and chlorine from widely different sources no such difference in the atomic weight has been observed.

THE DESCENDANTS OF THORIUM.

Thorium emanation	→ A	→ B	→ C ₂ C ₁	→ D
Average life . . . 54 sec.	0.14 s.	106 h. 60 m.		3 m.
Radiation . . . α-	α-	β-	α-	β-γ-rays
Estimated atomic weight } 220	216	212	212	208

With thorium, bismuth is said to be the hypothetical end product. The *onus probandi*, or burden of proof, of course rests with those who make the assertion. Inability to prove a statement is false is rarely of much value as evidence that the statement is true.

R. W. Gray (1913) reports that he found lead in some capillary tubes in which a minute quantity of radium emanation had been stored for four years. There are three possible sources of the lead: (a) The glass of the tubes contained 0.03 per cent. of lead; (b) The mercury used to seal the tubes contained a trace of lead; (c) the disintegration of the emanation, *à la* the emanation hypothesis. The glass of the tube was undoubtedly attacked by the emanation as shown by the discoloration, although no measurable amount of lead was obtained by digesting the powdered glass with water for some time.

§ 8. The Degradation of Uranium.

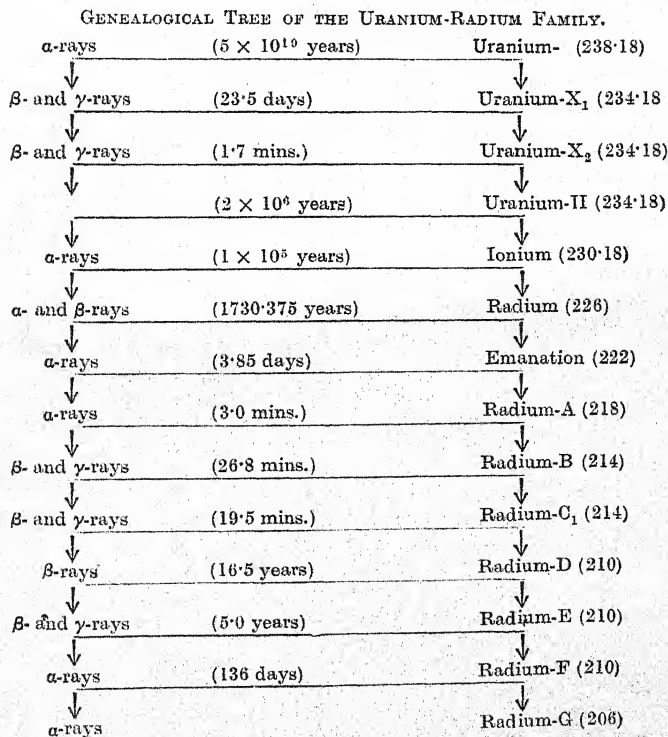
The habit of suspending judgment until the conclusion has been fully tested by varying the circumstances of the experiment, and by repeated accurate measurement, is a valuable habit to acquire.—G. F. FITZGERALD.

Under ordinary conditions, uranium and its salts—presumably freed from radium—give both α- and β-rays. If crystals of uranium nitrate, $\text{UO}_2(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$, be shaken with aqueous ether, the lower aqueous layer contains uranium which gives the α- and the β-rays; while the upper ethereal layer contains uranium which gives the α-radiations only. In order to distinguish these two varieties of uranium, the former is called uranium-X, and the latter simply "uranium" or "uranium proper." Uranium regenerates the normal quantity of uranium-X in from six to twelve months. Uranium-X is responsible for the β-rays of ordinary uranium; uranium proper gives only the α-rays. The extracted uranium-X loses its power of emitting the α-rays at the same rate as uranium proper regains it. A second quantity of uranium-X can be extracted from restored uranium, and so also a third and fourth extraction can be made; and, so far as we can guess, the extraction, restoration, and re-extraction can be repeated an indefinite number of times, that is, until all the uranium has been transformed into uranium-X. Hence, it seems impossible to avoid the inference that uranium is continuously and spontaneously decomposing into uranium-X and helium. In a similar manner it has been shown that uranium-X is itself breaking down into a radioactive solid which has been called ionium. Ionium bears some analogies with thorium.

Still further, it has also been proved that ionium is continuously and spontaneously passing into radium. Consequently, starting from uranium,

	Uranium \rightarrow Uranium-X \rightarrow Ionium \rightarrow Radium.			
Atomic weight	238.5	230.5	230.5	226.5
Average life	8,000,000,000 y.	35.5 d.	over 50,000 y.	2500 y.

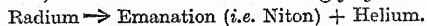
There is some evidence that what is here called uranium is itself a mixture of what have been called uranium-I and uranium-II, but that these substances are so much alike that no method of separation has yet been successful. There is also evidence that uranium-X furnishes two different products, uranium-X₁ and uranium-X₂. Without entering into further details, it can be said that there is a possibility that *uranium* is a *parent ancestor of radium*, and *radium* is a *parent ancestor of helium*, and, if inference be correct, that radium is the parent of lead. Summarizing these changes on a kind of family tree showing the nature of the rays emitted, the atomic weights of the products, and the average life :



Analogous tables have been compiled for the actinium and thorium families. Curiously, also, potassium and rubidium are slightly radioactive, for they emit β-rays, while the radioactivity of sodium has not been

detected. Therefore, between 30 and 40 radioactive bodies are assumed to exist; not half a dozen of these bodies have been isolated. The properties of those radioactive elements which have not been isolated have been deduced from their behaviour when mixed with large proportions of other known elements.

We have seen that the atomic weight of radium is 226.5 and of the emanation 222.5. It therefore follows that one atom of radium furnishes one atom of the emanation and an α -particle which, in turn, furnishes an atom of helium with an atomic weight 4. No other material product of the change is known, and the reaction is accordingly symbolized :



Similar assumptions have been employed in estimating the atomic weights of the other radioactive substances. The emission of β -rays (i.e. electrons) is supposed to produce no measurable change in the atomic weights, while the emission of α -rays (charged helium atoms) at each stage of the disintegration is supposed to reduce the atomic weight by 4.

At the beginning of 1913, several investigators stated that the expulsion of an α -particle by a radioactive element causes the residual product to shift its position two "places" in the periodic table in the direction of diminishing mass, so that the residual product is not in the next family, but in the next but one. Similarly, when an element gives off a β -particle, the product shifts its position in the opposite direction one "place" in a direction opposite to that for an α -ray change. Hence two changes attended by the emission of β -particles, and one by an α -particle would bring the product back to its original position in Mendeléeff's table. As A. S. Russell expressed the displacement rule: *whenever an α -particle is expelled by a radioactive element, the group in the table to which the resultant product belongs is either 2 units greater or two units less than that to which the parent belongs; and when an element gives off a β -particle, with or without the accompaniment of X-rays, the resultant product shifts its position so that it is one unit greater or one unit less than that to which the parent belongs.* This is illustrated with the radium family by Fig. 352. Similar tables have been compiled for the actinium and thorium families. Further, when any number of radioactive elements occupy one place in the periodic table, these elements cannot be separated from one another by any known chemical process. Thus, when mesothorium-I gives off two β -particles and one α -particle to form thorium-X, it is claimed that the two substances cannot be separated from one another in spite of the difference in their atomic weights; and they are probably spectroscopically indistinguishable. These non-separable elements are what F. Soddy called isotopic elements, or isotopes, e.g. ionium, thorium, and radiothorium are isotopes, and mesothorium is isotopic with radium (*vide supra*). The different forms of lead discovered by Soddy connected with radioactivity (1914) are also said to be isotopic, for they have the same atomic volumes and the same chemical properties; but F. A. Lindemann (1915) has shown that two elements of different atomic weight must differ either in their chemical or physical properties. The argument is based on the laws of thermo-dynamics, and he argues that the lead from thorite will probably have a melting point 1.54° higher than ordinary lead.

To summarize, the more salient facts are :

- (1) Radioactive substances are decomposing spontaneously.

- (2) The reaction is exothermal, but the speed of the decomposition is not affected by any known external condition ; the thermal value of the reaction, too, is more than a million times greater than that of any known chemical reaction.
- (3) The decomposition proceeds in a series of stages—consecutive reactions.
- (4) Helium is one ultimate product of the decomposition.
- (5) Three types of “radiant rays” are emitted at different stages of the decomposition.

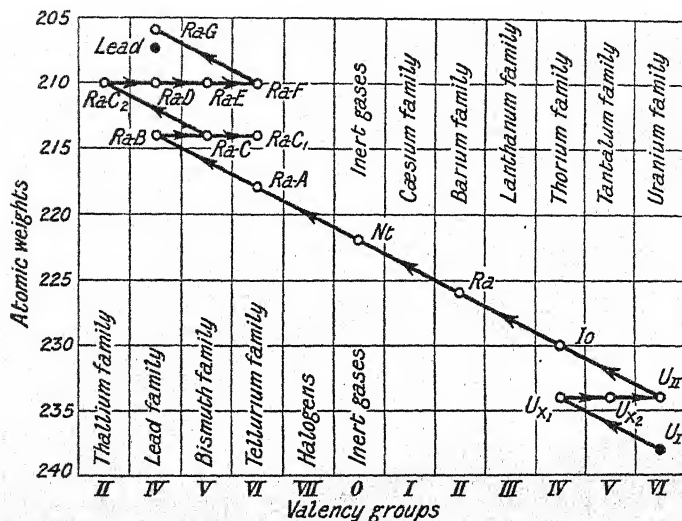


FIG. 352.—Arrangement of the Uranium-Radium Family in the Periodic Table.

The occurrence of radioactive substances.—Radioactive substances—chiefly thorium and radium—occur in many minerals in comparatively large amounts. In the uranium minerals, for example, the ratio Ra : U averages nearly 3.5×10^{-7} by weight. Radioactive substances are everywhere present in the atmosphere. It is estimated that a million parts of atmospheric air contain 0.06×10^{-12} parts of radium emanation, and 2×10^{-18} of thorium emanation. Hence, a charged electroscope in an exhausted glass vessel may retain its charge for months, but if air be admitted, the instrument will be discharged in a relatively short time—say, 24 or 36 hours. The emanations are found more particularly in the atmosphere over land than over the sea. Hence, it is inferred that the radioactivity of the atmosphere is derived from emanations from the land, and this is in agreement with the fact that the radioactivity of deep-sea water is greater near the bottom than near the surface. So far as observations have been made, every cubic centimetre of sea water averages about 0.017×10^{-12} gram of radium. This means that not less than the equivalent of 20,000 tons of radium occurs in the oceans of the earth. Radioactive substances are widely distributed in the solid crust of the earth;

they are found in all kinds of rocks, clays, soils, etc. One estimate indicates that an average of 1.4×10^{-12} gm. of radium is present per gram of rock. The water from deep-seated springs and wells has usually a marked radioactivity, and it has been stated that the curative properties of certain mineral waters—e.g. those of Buxton, Bath, Wiesbaden, Bohemia, etc.—are due to their radioactivity. This statement is mere guessing, because so little is known about the physiological action of the radiations from radioactive materials. The fact that radioactive substances constantly emit heat, and since the emanations are ubiquitous in the earth's crust, it follows that the heating effects of these substances must play an important part in maintaining the heat of the earth, and must have profoundly modified the rate of cooling of the earth in past ages.

§ 9. The Mutability of the Elements ; and the Disintegration of Atoms.

To inquire whether the metals be capable of being decomposed and composed is a grand object of true philosophy.—H. DAVY (1811).

Radioactivity is the least manageable of natural processes. It will not be hurried or controlled. Nature keeps the management of this particular department in her own hands. Man views the phenomenon with hungry eyes, but his interference is barred out. He can only look on in wonder while it deploys its irresistible unknown forces.—ANON. (1907).

About 1902, many hypotheses were suggested to explain the phenomena associated with radioactivity ; R. Meldola, for example, suggested the helide hypothesis of radioactivity. In this, the radioactive elements are supposed to be compounds of helium with other elements—helides—and these compounds are further supposed to be undergoing a gradual spontaneous decomposition into simpler substances. This hypothesis has been abandoned by R. Meldola ; it has received very little consideration, because attention has been mainly focused upon a second hypothesis, which runs somewhat as follows : Ordinary atoms are supposed to be small intricate systems of electrons, linked together by forces of tremendous power. The properties of the different elementary atoms are determined by the number and configuration of the intra-atomic electrons. Radioactivity is an atomic property, and it is an effect of the instability of certain atomic systems. The disintegration of the unstable atoms is marked by the emission of rays. The radioactive elements are therefore unstable, and are continually and spontaneously changing by numerous intermediate stages into more stable elements. This hypothesis is called Rutherford and Soddy's theory of the disintegration of the atoms, because they established its claim to serious consideration, and have done valuable work with its aid. Their hypothesis is orthodox and fashionable. If this hypothesis should survive that struggle for existence which all neoteric hypotheses must undergo, then radioactivity will be cited as proof of the devolution of the elements. Astro-spectral observations, p. 1006, leave little room for doubt that in the "cooling stars" a process of evolution of the elements is in progress.

It is further assumed that the radioactive elements are not unique among the elements in containing abnormal stores of internal energy, but—excluding potassium and possibly rubidium—the other elements are either immutable or else they are changing so slowly that no signs of mutation have yet been detected. According to the atom disintegration

hypothesis or radioactivity Nature is continually changing the elements with the largest atomic weights such as uranium (238.5) and thorium (232.4) into simpler elements. The latter, in turn, are said to be stable simply because no signs of radioactivity have yet been detected. It is possible that if ever elements existed on earth with larger atomic weights, and by inference, with more complex atoms, they have all degraded into simpler forms, and are now probably extinct elements, hence also, it might be inferred that *the most widely diffused elements have small atomic weights*. The gaps which appear in Mendeléeff's table (p. 992) also appear significant. The elements with the smallest atomic weights, and those which are found in greatest abundance on the earth—hydrogen, helium, calcium, oxygen, sodium, silicon, etc.—are usually considered to be the most stable, and to contain least infra-atomic energy. Hydrogen and helium, occurring in the hottest stars, are supposed to have a tendency to form aggregates, and pass into common terrestrial elements during the cooling of the hot stars. It seems as if uranium and thorium must have been exposed to peculiar conditions—possibly of pressure and temperature—whereby they were elaborated beyond the limits of stability, and absorbed stores of energy which are now being slowly released because the conditions necessary for their stability no longer obtain.

It might be asked why the comparatively conspicuous self-destructive activities of radium have not led to its extinction long ago? Rutherford estimates that the radium now on earth will be disintegrated and the whole virtually extinct in about 25,000 years. There can thus be little doubt that if there had not been a continuous source of supply, radium would have been an extinct element long ago. The decay of the heaviest known element uranium is so extraordinarily slow that it can just be detected, and a rough estimate made of its life—8,000,000,000 years—as indicated above.

The mode of evolution of the elements hypothesized in the attempt to co-ordinate the results of the spectroscopic study of stars and nebulae, seems to be supported by a mass of cumulative evidence (p. 1001), and to be inherently probable. It remains to find an adequate explanation to account for the vast stores of energy available in the hotter stars. Here again we are confronted with mystery profound (p. 149).

The distribution of the elements.—The reasons for thinking that the supplies of radium are continually renewed turn on the facts:

1. Radium and uranium always occur together, and the two elements are not sufficiently common for this to be due to mere chance; and

2. The proportion of radium to uranium in the uraniferous minerals is almost invariable—1 : 35,000,000. This approximate constancy is clearly the result of an equilibrium between production and decay. The supply of radium is regulated by its relative rates of formation and degradation; and when the speeds of the opposing reactions are balanced, the ratio radium : uranium must be constant. It is interesting to observe in this connection that a very small amount of helium is always found occluded in uraniferous minerals. Assuming that no helium escapes, the small amount found in a given rock will be a measure of the time which has elapsed since the birth of that particular sample, but this gas must be constantly leaking into the atmosphere, and, consequently, the "age" so computed will be a minimum age of the mineral, for the mineral may be

older, but not younger than the age so computed. Hence, by determining the relative amounts of uranium and helium in a mineral, its minimum age can be estimated. In this way, Strutt estimated that it requires eleven million years to produce one c.c. of helium per gram of uranium. Lead also has been detected in over forty uraniferous minerals; and in many of these cases lead does not occur near the uranium deposit. Hence it is considered unlikely that the lead has been deposited about the uranium by subterranean streams. The general notion is that lead is the "final" product of the degradation of uranium (p. 1029), in spite of the fact that the change is too slow to be observed directly, and the evidence very flimsy.

We can now see a possible explanation for the "inseparable companionship" of many elements emphasized on p. 1002. The "occurrence" or "distribution" of the elements over the earth appears to be an effect of an important genetic law. The approximate uniform quantities of many of the rarer metals in different parts of the earth indicate that these elements are degradation products of more complex elements; and that they, in turn, will probably be degraded into simpler products. The relative speeds of these slow changes determine the amount of each element which can be present on the earth at any given time.

The student may very properly think that *a modicum of fact is here entangled with an abnormal amount of speculation*, especially when it is remembered that the experiments have been made upon very minute quantities of material. At first sight, it does appear as if we have developed what A. Smithells humorously called "a chemistry of phantoms." Thanks, however, to the extraordinary delicacy of the electrometer and of the spectroscope, there is no doubt about the facts, even though but extremely minute quantities of radium are available for experiments. The argument converges on the assumption that radioactivity is an atomic property; this hypothesis, in turn, is mainly based on the indifference of the speed of radioactive changes to external conditions of temperature, pressure, etc. It is therefore pertinent to inquire into the validity of the alternative hypothesis, and ask:

Are the "radioactive elements" really elements, or are they compounds containing helium? As already hinted, the dogma that radium is an element is not so firmly established that there are no reasonable grounds for the exercise of some Cartesian doubt (p. 903), for, said E. W. Morley (1909), he is wise whose assertions regard the possibility of finding at some time evidence to the contrary. E. Rutherford (1909) has said that "since in a large number of cases the transformation of the atoms is accompanied by one or more charged atoms of helium, it is difficult to avoid the conclusion that the atoms of the radioactive elements are built up, in part at least, of helium atoms." As an alternative to Rutherford's atom-disintegration hypothesis, and as a corollary to the inference that the molecules of the argon-helium family are really polyatomic (pp. 689, 694), R. Meldola (1903) suggested that the radioactive elements are really compounds of an active form of helium in the same sense that nitrogen chloride may be said to contain the atomic or active form of nitrogen. Ordinary helium, like ordinary nitrogen, is characterized by great chemical inactivity. If this hypothesis be valid, it must be supported by a formidable list of unique hypotheses, for it must be assumed further that (1) the alleged compound—helide—is spontaneously decom-

posing; (2) abnormally large amounts of energy are set free during the decomposition; (3) the emission of radiations accompanies the change; (4) the speed of the decomposition is not affected by any known external conditions; and (5) in opposition to a little circumstantial evidence, that the helium molecule is a complex of atoms, and not monatomic. *Otherwise*, adds H. E. Armstrong (1913); the decomposition of radium regarded as a compound of atomic helium is no more remarkable than that of liquid ozone, or of nitrogen chloride! Armstrong continues:

The atoms of helium and of the allied inert gases are gifted with intense activity far beyond anything we know of; it may well be that when such atoms enter into combination, either with one another or with other elements, the amount of energy set free is very great, and that when they combine with other materials, they may produce changes in properties very different from and far more profound than those we know of at present.

The plausibility of the helide argument turns on this: The greater the amount of hypothetical energy assumed to be required to break down the hypothetical helide molecule into its supposed atoms, the greater the appearance of probability of the argument that radioactive phenomena are chemical in kind. Questions like these have to be treated more by instinct (or prejudice) than by logic; and those who do the work must use what hypotheses they find most fruitful.

H. S. Sheldon (1913) has emphasized the fact that the indifference of radioactive changes to temperature and other physical conditions must be a relative phenomenon, and a consequence of the limited range of our resources. The highest temperatures of our laboratories—3000°—are feeble when contrasted with those ten times hotter, which prevail in the colossal furnaces revealed by stellar spectra of the hotter stars. Consequently, the indifference of radioactive changes to external conditions cannot be accepted as absolute. Suppose, he adds, that electricity were unknown and it was only possible to attain variations of temperatures of a few degrees in our laboratories, then a large number of so-called compounds would be classed as elements, and the slow decomposition of many substances with the evolution of heat would appear as marvellous sources of energy, as unaccountable as radioactive changes are to-day.

§ 10. The Effect of the Discovery of Radioactivity on the Definitions of Element and Atom.

We must fight against the soporific influence of sham definitions. Especially must we guard against ever allowing them to be in the way of an inquiry into facts.—A. SIDGWICK.

Naturally, the student of chemistry may be somewhat disconcerted with this apparent attack on what appear to be the essential principles of chemistry outlined in the earlier part of this text-book. At first sight, it seems as if we must say "good-bye to the equations of chemistry," because, if the truth about the suspected disruption of atoms be ever vindicated, it appears that the fundamental concepts—atom, element, persistence of weight, etc.—must be revised in order to make them describe the facts.

The elements.—The conception of an element, given on p. 24, has long held an honoured place in chemical text-books; and, with this before us, it certainly appears illogical to apply the term to a substance which can be resolved into two or more simpler forms of matter. Any substance

which can furnish two or more different elements has certainly forfeited its place in the list of elements. Some try to evade the difficulty by assuming that there is an agreement among chemists to recognize a substance as an element which, under proper conditions, exhibits a spectrum showing characteristic lines possessed by no other element, and possesses a definite combining weight. Of course, we are at liberty to change our definitions, but, as W. Bancroft (1912) has said, the only advantage of the new definition is that it enables chemists to say that they have decomposed an element. S. Curie (1911) stated definitely that radium is not a compound of helium, but the only way she can make radium an element is by changing the old definition so that it shall include radium. The definition on p. 25, however, is elastic enough to cover cases of transmutation, for, if an element containing but *one kind of matter* suffered an atomic cataclysm, it might furnish two or more different kinds of matter each of which is an element because it contains but one kind of matter. It would probably lead to clearer thinking if a special term were invented. For example, some one has suggested that the elements be called *chemical primaries*. We are always loath to multiply definitions, and would much rather condense a number of definitions into one; here, however, there seems a special need for distinguishing terms. As already emphasized, the definition must not be taken to imply that the elements are absolutely immutable, although, so far as our present knowledge goes, they are both immutable, in all chemical changes, and primitive. The isotopes of an element must also be regarded as different elements each having a characteristic atomic weight, rate of diffusion, specific gravity, and other specific properties when we know more about them.

The atoms.—Paradoxical as it may seem, the hypothesis that the atom of an element is a most intricate bit of mechanism, a complex aggregate of parts liable to disruption, is now generally accepted. This, however, does not affect the time-honoured definition of an atom indicated on p. 50. **The atom still remains a veritable unit indivisible in chemical reactions.** Had the facts, speculations, and theories discussed in these concluding chapters been treated at the beginning of our study, that would not have altered our mode of presenting the facts of material chemistry.

The law of persistence of weight.—Suppose an atomic convulsion or cataclysm were to occur so that the complex system of electrons which is supposed to form an atom were to break up into simpler parts; suppose further some of the electrons grouped themselves into helium, and the others into some other substance; and let us also assume, for the moment, that some electrons simultaneously escape and are merged into the æther of space, it is then conceivable that there will be an apparent loss of weight. Consequently, while *the law of persistence of weight holds good with chemical reactions in which the atoms remain intact*, it is quite conceivable that an apparent loss in weight might occur during a radioactive change (p. 20). If the products of the disintegration of the atom have mass, it might be inferred that the absolute mass still remains constant, although, if radium be an element, the invariability of mass or weight can no longer be referred back to the constancy of the atom. This visionary phenomenon has been described in order to emphasize the need for care in building rigid, nonplastic concepts and definitions from negative results based on the uncontradicted experience of mankind, p. 134.

§ 11. The Transmutation of the Metals—Alchemy.

Would to God all men might become adepts in our art, for then gold, the great idol of mankind, would lose its value and we should prize it only for its scientific teaching.—E. PHILALETES (c. 1623).

Alchemy appears to have been a medieval system of philosophy, and it sought to demonstrate the validity of its doctrines concerning the cosmos by transmuting the baser metals into gold. The following excerpts from alchemical writings will perhaps make clear their ideas on this subject:

Gold is the most perfect of metals because in it Nature has finished her work.—ROGER BACON (1214—1294).

The metals are all essentially identical, they differ only in form. Form brings out the accidental causes which the experimenter must try to discover and remove.—ALBERTUS MAGNUS (1193—1284).

If by any reason the superfluous matter could be organically removed from the baser metals, they would become gold and silver. Our art only arrogates to itself the power of developing, through the removal of all defects and superfluities, the golden nature which the baser metals possess.—E. PHILALETES (c. 1623).

The mystic gnome who laboured in order to quicken the growth of the metals in mines was a creature of the superstition of the childhood of man. The, perhaps, more philosophical alchemists believed that the perfecting of the metals occurred spontaneously in the "bowels of the earth," and Pliny tells us that exhausted mines have been closed down to enable the metals to fructify, and so be again profitably worked in a few years' time. They also believed that Nature aimed at the production of gold in mines, and that when she is hindered in her design, the so-called imperfect metals appeared. The baser metals were hence called "diseased gold"; mercury was "ailing silver"; copper, iron, lead, and tin were "lepers" which, when cured of their leprosy, would become gold. The alchemists sought to find some means which would hasten the slow natural change so that the transmutation could be conducted in a much shorter time. "It is this means," said Roger Bacon, "which the alchemists indifferently called the elixir, the philosopher's stone," etc.

The alchemists wrote in a language we do not now understand. They seem to have associated mystic extravagances with their operations, and to have described their processes in obscure ambiguous jargon, with the result that their readers were "stunned," as Boyle expressed it, "with dark and empty words." No wonder that J. Potanus (1520) complained that after travelling through many countries to examine the claims of the adepts, he found "many deceivers, but no true philosophers"; and N. Lemery (1675) that "they professed an art the beginning of which was deceit, the progress of which was falsehood, and the end beggary." Alchemy thus fell into disrepute, for it seemed as if its claims could be established only by chicanery and fraud. Some of the more honest believers explained their failure by asserting that "the art of making gold is beyond the reach of human capacity, and it is made known by God to those alone whom He favours, and who are called adepts." The majority of alchemists, however, did seek to make gold cheaply with the sole object of gaining "untold wealth." Failure or delusion was inevitable. Accordingly, the alchemist often misrepresented the truth and degenerated into a charlatan and impostor, pretending with vulgar frauds, that he had

succeeded "in order," says M. M. P. Muir, "that he might really make gold by cheating other people," or else gain notoriety.

The idea of transmutation appears to have come from the Egyptians and Chaldeans, but its real origin is lost in remote antiquity. M. Berthelot (1885) considers that the idea did not originate from the philosophical views of the ancients on the unity of matter as is stated above, but rather from the attempts of the goldsmiths to make fraudulent substitutes for the precious metals.

It is easy to understand how the belief that the base metals could be converted into gold dominated ancient and medieval chemistry. Facts were cited in its favour. The production of beads of silver and gold by the cupellation of metallic lead, and the reduction of metallic ores furnished direct evidence of the metamorphosis of the metals. Again, iron utensils in copper mines became coated with red copper when left in contact with the "mine water," so that the iron was seemingly transmuted into copper; similarly, the formation of white and yellow alloys by mixing copper and certain earths, seemed with the then imperfect knowledge, ample proof of transmutation. The dogma of transmutation thus appeared eminently plausible; it ran counter to no known laws of nature; it rested upon no extravagant assumptions; and it was sanctioned by the highest authorities. The immense labour which must have been expended in the fruitless pursuit of this chemical chimera by the alchemists is appalling. The quest was virtually abandoned with the advent of Lavoisier's balance.

I. Remsen (1903) has emphasized the fact that the alchemists were the working chemists of their day, and that they laid the foundations of experimental science. He further develops the idea that there is a life after the death of a good doctrine. The phlogiston theory lived in the form of the modern doctrine of free energy, and the idea that there is a relationship between the elements is what Remsen calls the spiritual part of alchemy which lives though alchemy is dead.

The alchemist's dream of transmutation is little if any nearer realization to-day than it was a thousand years ago, for no one has yet really succeeded in transmuting one chemical element into another other than by speculative argument. There is no unimpeachable evidence of a single transmutation of one element into another pre-determined by man. In the words of S. Curie (1912): "*on ne peut considérer qu'il y a pas encore actuellement de raisons suffisantes pour admettre que la formation de certains éléments puisse être provoquée à volonté en présence de corps radioactifs.*" True enough, a few radioactive elements—radium, actinium, polonium, uranium, and thorium—seem to have been discovered in nature, and they are usually stated to be changing spontaneously from one elemental form to another; but no process known to man is able to accelerate or retard, stop or start the metamorphosis. No element has yet been broken down into a simpler substance by a process controllable by man. "Nature to be conquered must be obeyed" (Francis Bacon).

The alleged transmutation of copper into lithium and sodium by A. T. Cameron and W. Ramsay (1908) has been denied by S. Curie and Gleditsch (1908) and by E. P. Perman (1908)—the lithium and sodium were derived from the vessels used in the work; the production of neon from radium emanation by W. Ramsay (1907) and W. Ramsay and A. T. Cameron (1908) has been denied by E. Rutherford and T. Royds (1908)—the neon was derived from the air which had not been excluded from the apparatus; and the formation of carbon dioxide

by the action of radium emanations on solutions of thorium and zirconium by W. Ramsay and F. L. Usher (1909) has been called in question by E. Rutherford (1913)—the carbon appears to have been derived from the grease used in lubricating the stopcocks. The alleged transmutation of hydrogen into neon, by W. Ramsay (1913) and J. N. Collie and H. Patterson (1913), by the action of a stream of cathode rays on hydrogen is considered by J. J. Thompson (1913) to be a malinference, since the neon is thought to be derived from that originally occluded by the electrodes, or glass vessel, and which is expelled by the bombardment of the cathode rays, but which cannot be removed by the mere application of heat. R. J. Strutt (1914) and T. R. Merton (1914) could not verify the alleged conversion of hydrogen into neon.

The last paragraph but one may want modifying some day. E. Rutherford counted the number of scintillations per second produced by the α -rays from a given source on a zinc sulphide screen *in vacuo*, and then introduced dried carbon dioxide; the number of scintillations per second decreased in accord with the known stopping power of that gas; when nitrogen was introduced the number of scintillations per second increased. It was therefore inferred that the nitrogen atom is disrupted by being bombarded with α -rays, and that one product of the action is a swiftly-impeled hydrogen atom. The deduction was confirmed by estimates of the mass from the magnetic deflection of the particles. E. Rutherford further estimated that in the case of nitrogen only one α -particle in 300,000 succeeds in getting near enough to the nucleus to liberate a swift nitrogen atom with sufficient velocity for it to be detected; and if the entire α -radiation of a gram of radium were absorbed by nitrogen, it would generate about 0.0000005 c.c. of hydrogen per year. He added: It may be possible that the collision of an α -particle is effective in liberating hydrogen from the nucleus without necessarily giving it sufficient velocity to be detected by scintillations. If this should prove the case, the amount of disintegration may be much greater than this estimate. It would be a bold chemist who would claim to be able to clean his products free from hydrogen or its compounds to anything approaching the implied degree of purity. E. Rutherford and J. Chadwick, however, stated that while bombarding α -particles have a range of 7 cm. the swiftly-moving hydrogen atoms obtained from hydrogen gas and hydrogen compounds have a maximum range of 29 cm. in air, while those from nitrogen have a range of 40 cm. This is taken as evidence that the swiftly-moving particles cannot arise from the contamination of nitrogen with hydrogen. They also found that the long-range hydrogen particles were furnished by fluorine, aluminium, phosphorus, boron, and sodium; but the numbers from the last two elements were less than from the other three. The following elements were bombarded, but very little if any effect was observed: lithium, beryllium, carbon, oxygen, magnesium, silicon, sulphur, chlorine, potassium, calcium, titanium, manganese, iron, copper, tin and gold.

Intra-atomic energy.—The facts previously indicated show that not far from 2,000,000,000 cal. of heat are evolved during the degradation of one gram of radium. This is a quarter of a million times greater than is evolved by the combustion of a similar weight of coal. Hence it is inferred, from the atomic disintegration hypothesis of radioactivity, that the atoms of the radioactive elements, and probably also of other elements, have tremendous stores of potential energy, far greater than is developed during ordinary chemical reactions. The rate of degradation of the energy of the radioactive elements is comparatively slow, and is

not available for doing useful work. The rate of evolution cannot be influenced by any known conditions and consequently the transmutation of the elements involves the discovery of methods of controlling these tremendous supplies of energy. Just as the application of a large quantity of electrical energy concentrated at the ends of a pair of platinum wires enabled H. Davy (1808) to decompose the alkalis, so W. Ostwald, W. Ramsay, and others infer: *if ever one stable element is transmuted into another element, a large quantity of energy in a highly concentrated condition will be required.* This is quite in harmony with the alleged dissociation of the elements in the hotter stars (*q.v.*) where but a few elements are present, and where the temperature has been estimated at 25,000°. The 3000°–4000° obtained in some electrical furnaces appear but puny in comparison with the tremendous natural powers present in the hotter stars—p. 449.

It has been pointed out that the formation of, say, gold from a metal atomically lighter, say tin, would require the expenditure of so much energy that even if the transformation were accomplished, it could not be a successful commercial process for the production of gold. On the other hand, the formation of gold from an atomically heavier metal, say lead, would liberate such an enormous amount of energy that the gold would be but an insignificant by-product, for the energy liberated during the process would have an enormously greater value than the metal.

§ 12. The Creation and Annihilation of Matter.

Even if we resolve all matter into one kind, that kind will need explaining
And so on for ever and ever deeper and deeper into the pit at whose bottom
truth lies, without ever reaching it. For the pit is bottomless.—O. HEAVI-
SIDE.

Things which are seen were not made of things which do appear.—ST. PAUL.
Æther is the parent of all things.—LUCRETIVS (c. 60 B.C.).

In mechanics, the definition of matter is based upon Newton's first law of motion—the law of inertia—where matter is defined as that which requires the expenditure of an external force to change its state of motion. Otherwise expressed, inertia, or helplessness, is a characteristic of every form of matter. No material thing can of itself change its own state of motion, for an external influence is required before such a change can take place. If it be admitted that any entity which requires the application of a force before it can change its state of motion is said to be a form of matter, an electron in motion must be a form of matter, because it requires the application of a force to change its state of motion.

The inertia of matter.—This definition is also reversed, and force is defined to be that influence which is required to change the velocity of any material body, and it is measured in terms of the dyne as unit. A dyne is that force which applied to a mass of one gram during one second imparts to it a velocity of one centimetre per second, or which changes the velocity of the body one centimetre per second when acting for one second. These ideas can be expressed in another form. The inertia of a body is that property of matter which resists change of motion, and it is measured in terms of the force required to produce a change of one centimetre per second in the velocity of the body. Accordingly, *when the action of a force on two bodies produces the same change of velocity per second, their inertia are said to be equal.*

The mass of matter.—A body falling from a height down to the

earth's surface—in *vacuo* so as not to be influenced by the resistance of the air—gains in velocity, say, g cm. per second in each second. If the mass of the body be m , then the force pulling that body down will be measured by the product mg . This is the weight of a body; hence the weight of a body is g times its mass. The weight w of a body measures the force by which it is attracted to the earth's surface, whereas mass refers to the quantity of matter m in a body, and is independent of gravitation or weight. Experiments have shown that all material bodies have the same numerical value for g , namely, 981 cm. per second when acting for one second (latitude 45° , and at sea-level). Accordingly, $w = 981m$; or the weight of one gram of matter is 981 dynes; or the weight of a body is 981 times its mass; or the mass of a body is $\frac{1}{981}$ of its weight—under standard conditions. The masses of two bodies in the same place must therefore be proportional to their weights. *If in the same locality the action of gravity on two bodies produces the same change of velocity per second, their masses are said to be equal.* This conclusion is in harmony with that deduced in the preceding paragraph, and, accordingly, it has been inferred that mass and inertia are identical; meaning that the definitions of inertia and of mass, in the nomenclature of mechanics, are not mutually exclusive.

Electromagnetic mass.—When charges of electricity are set in motion, they act like electric currents, and set up magnetic fields which oppose the motion—Lenz's law. Motion against the induction effects requires an expenditure of energy, and therefore a moving charged particle will appear to have an added electromagnetic inertia in virtue of its constituent charges. This added inertia will be proportional to the total electrical energy of the charges. H. Kaufmann (1906) measured the value of the numerical ratio e/m , where e represents the electric charge, and m the inertia or the mass, for electrons travelling at different speeds, and found—

Velocity $\times 10^{-10}$	2.83,	2.72,	2.59,	2.48,	2.36 units.
$e/m \times 10^{-10}$	0.83,	0.77,	0.97,	1.17,	1.31 units.

Hence, either the value of m increases or the charge e decreases with increasing velocity, for the greater the velocity the smaller the numerical value of the ratio of the electric charge to the mass. Rightly or wrongly, it is assumed that the charge on the particles remains constant and invariable, and accordingly, it follows that the (electromagnetic) mass of the electrons is not constant, but increases rapidly as the velocity is augmented. Further, if any part of the mass of an electron is ordinary mechanical mass, it must be very small in comparison with that which is of electrical origin, since the electrical inertia of a body depends upon its velocity and approaches infinity when the velocity of the body approaches that of light. This variation in the apparent inertia or mass of a body is the same as if the electrical mass existed alone, and the material mass were virtually zero. Consequently it has been inferred that (1) the electrons do not possess a material mass in the ordinary sense of the word; and (2) the electrons have no mass other than that which is derived from their motion and electrical charge. If the real mass of an atom is the sum of the positive and negative electrons, and the latter have no material mass, it might be inferred that the positive electrons or the mass of an atom would still remain a constant. H. A. Lorentz, however, has shown that in all probability the masses of all particles will be affected by their translational velocity to the same degree as the electromagnetic mass of the electrons.

Again, it is assumed that electricity is a phenomenon of the æther and that the atom is built entirely of electrons; consequently, as J. Larmor (1900) expressed it: atoms are forms of æthereal strain, or the material atom is formed entirely of æther, and has no material substratum. Further, the electrons are supposed to be minute electrified strains, squirts, vortices, swirls, eddies, or whirlpools in the æther. The intangible, imponderable, all-pervading æther is supposed to be susceptible to the solicitations of gravity, as matter, only when it is stirred into innumerable swirls, which are held together by powerful cohesive forces. Hence, it is said: *æther is the mother of matter*; for in the æther—

Wrapt in mystic silences and glooms,
The slum'ring secrets of creation lie.

Some even claim to have created material atoms from immaterial æther, but the claims are somewhat questionable. The stages in the alleged genesis and destruction of matter can thus be symbolized:

Material	Immaterial
Matter <i>en masse</i> \rightleftharpoons Molecules \rightleftharpoons Atoms	\rightleftharpoons Electrons \rightleftharpoons Æther.
Dematerialization of matter \rightleftharpoons Materialization of æther.	

When, therefore, it is assumed that material electrons have been resolved into primitive æther by relieving the strains, or stilling the swirls, it is at the same time assumed that matter can be deprived of cosmical existence, and transmogrified into Nirvanian nothingness. With the annihilation of ponderable matter, the law of the conservation of matter disappears, and apparently the only constant in the universe is nothingness!

The argument turns on the nature of the electrons and of the æther. The æther is assumed to be immaterial, or, as Pythagoras (c. 500 B.C.) expressed it, "a celestial substance free from all perceptible matter." As a matter of fact, the æthereal plenum cannot be compared with any known thing. It is utterly beyond the range of our sense perceptions, and as G. le Bon (1907) has said, we are related to the æther much as a man born deaf is related to music, or a man born blind is related to colour; accordingly, no analogy can make such men understand what is a sound or a colour. In the words of T. A. Edison (1893): "As for the æther which speculative science supposes to exist, I know nothing about it."

§ 13. Thomson's Corpuscular or Electronic Hypothesis of Matter.

If we be curious to know what matter is, we plunge at once into that deep which surrounds us on every side, and which never yet was fathomed by human intellect.—J. F. DANIELL.

Liability to error is the price we have to pay for forward movement.—A. SIDGWICK.

No difference other than velocity of translatory motion has been detected in the properties of negative electrons when produced in many different ways, and from many different gases; and since the mass of a negative electron (corpuscle) is less than that of any known atom, the corpuscle must be a constituent of many different substances; and the atoms of these substances consequently must have something in common. This suggests the idea that the atoms of the chemical elements are built of simpler components, the electrons have thus been regarded as the ultimate sub-atoms or the *Uratomen* of which matter is essentially

composed. Since the electron is a constituent of atoms, J. J. Thomson considers that "it is natural to regard the electron as a constituent of the primordial system." J. J. Thomson's electronic theory of matter is one of the unitary theories of matter. It assumes that each atom of any element consists of a large number of electrons, all electrified negatively, and held together by positive electricity equivalent in amount to the sum of the negative charges of all the electrons so as to produce an electrically neutral atom. Thomson also accepts the atom-disintegration hypothesis of radioactivity.

The distribution of a number of negatively charged particles in a sphere of uniform density has been investigated mathematically by J. J. Thomson (1904). Five is the greatest number of electrons which can be in equilibrium in a single ring. But if other electrons be placed within the ring, a larger number can be maintained in equilibrium in one ring. Thus a ring containing six electrons would not alone be stable, but if a seventh electron be placed within the hexoidal unstable ring, the system will become stable. A greater number of electrons will arrange themselves in series of concentric rings.

Mayer's floating magnets.—The idea was neatly demonstrated by repeating an old experiment due to A. M. Mayer (1878-9). Small uniformly magnetized needles were thrust through discs of cork, and floated on water so that the negative poles of all the needles floated above the surface of the water while the positive poles were submerged. These needles arrange themselves like J. J. Thomson's imaginary corpuscles when a positively charged magnetic pole is suspended a little above the surface of the water. The diagram, Fig. 353, shows that a group of four needles arrange themselves in the water at the four corners of a square; if another needle be thrown into the water, the five needles take up positions at the corners of a pentagon; if another needle be thrown in, five needles form a pentagon as before, but the sixth needle goes to the centre of the pentagon; if another needle be introduced, six arrange themselves at the corners of a hexagon, and the remaining needle goes to the centre.

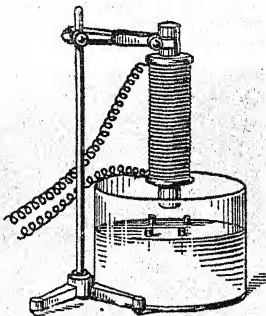


FIG. 353.—Mayer's Floating Magnets.

Thus, a ring of six needles is unstable if hollow, but it is stable when another is placed inside. This is an important principle in the formation of stable systems of negatively charged electrons. It is obvious that for stable equilibrium the structure must be substantial; a system with a large number of electrons on the outside, and none within, will be unstable.

There are important differences between Mayer's floating magnets, Thomson's systems of concentric rings of corpuscles, and the corpuscles in a real atom, because the two former are supposed to move in one plane (as on the surface of the water), whereas the constituent corpuscles of an atom could no doubt move in any direction in space; and the conditions of stability of systems of electrically charged corpuscles must be altogether different from those obtaining in Mayer's experiment.

The architecture of the atom.—According to Lord Kelvin's hypothesis (1902) each atom is a remarkable constellation of electrons ranged in a system of concentric rings all assembled within a sphere of uniform positive electrification. Although the mass of the negative electron seems to be $\frac{1}{1700}$ of that of the hydrogen atom, it cannot be assumed that the hydrogen atom contains about 1700 negative electrons, unless the mass of the positive electron be ignored. The surface of the imaginary sphere may be regarded as the limiting surface of the atom. The electrons are further assumed to be in rapid orbital motion about the centre of the sphere, and they are therefore driven outwards by a definite centrifugal force. They are subject to the mutual repulsion of the negative electrons and the attraction of the total positive charge. Positive electricity never appears apart from matter, so that it is always associated with the atom itself. Some very interesting investigations have recently been made on the nature of the positive electrons. In order to explain spectral phenomena, Nagaoka (1904) assumed that the positive charge is concentrated at a point in the centre of the atom, and the charge is not then uniformly distributed over the sphere. As H. Poincaré expressed the idea: "The neutral atom may be considered to be composed of an immovable principle portion positively charged round which move, like satellites round a planet, several negative electrons of very inferior mass." Fig. 354 gives but a crude notion of what is meant by concentric rings of corpuscles in rapid orbital motion about a central positively charged electron. The atom has thus been compared with the planet Saturn and its rings. This

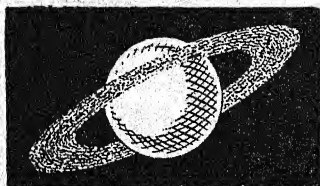


FIG. 354.—Diagrammatic Representation of an early Form of the Corpuscular Ring Hypothesis of the Constitution of an Atom.

intricate mechanism is supposed to represent the inner structure of the chemist's atom. The atoms, in turn, are supposed to be disposed in the molecule in an analogous manner. Consequently, the relation of the electrons to the whole molecule must be somewhat complex. The apparent hardness and impenetrability of the atom, so well emphasized by Newton, p. 47, now appears to be a kind of gyrostatic equilibrium due to the rapidity of motion of the component electrons.

We are told that a soft wax candle fired from a gun can be shot through a deal board, and that the speed of the jets of water in hydraulic mining in California was so great that a man could not strike an axe into the water which had just left the nozzle. It is quite possible therefore that an atom may appear to be rigid owing to the rapid motion of the constituent electrons. "Hardness," it is said, "is merely softness in rapid motion."

The Saturnian atom.—One reason for postulating the Saturnian atom as a basis for investigation is rather interesting. J. Plateau spun little spheres of oil in a mixture of alcohol and water so that they rotated while suspended in the liquid. During their rotation, the little spheres flattened at their poles and bulged at their equators. When the speed of rotation of a sphere is great enough, rings of oil were thrown off. The rings broke and coalesced to form little spherules which rotated on their own axes and revolved round the parent sphere. This remarkable experiment has been used for illustrating a possible mode of formation of a solar

system like ours from a contracting spinning nebula. The Saturnian atom is supposed to be the figure of equilibrium of a mass of rotating electrons in a sphere of positive electrification. It is not difficult to understand how the "figure of equilibrium" of a rotating fluid cannot be spherical, for gravity is progressively neutralized in passing from the axes of rotation (the poles) towards the equator, where the mass is specifically lighter. The faster the spin the more oblate the spheroid, and in the extreme case, with an increasing speed of gyration, the spheroid must become either a flattened disc, or else, at some critical velocity, there must be an interruption in the process of flattening, and the spheroid must alter fundamentally in shape or break into pieces. H. Poincaré (1885), G. H. Darwin (1879-81), and J. H. Jeans (1904) have investigated the subject mathematically.

Rutherford's atom.—E. Rutherford showed that when the α -particles approach the nucleus of an atom, their paths are deflected from the rectilinear course. This can be illustrated by Fig. 355. The positively charged α -particles are supposed to be travelling from left to right in the vicinity of the positively charged nucleus of an atom, +, the repulsive effects produce a scattering of the rays as indicated diagrammatically by the arrows in the diagram. In order to explain the scattering of the α -rays by matter, E. Rutherford (1911) assumed that the atom contains (i) a massive nucleus no larger than 10^{-16} cm. in diameter and charged with positive electricity. (ii) A cluster of negative electrons revolving about the central nucleus in orbital motion. (iii) An outer group containing but a few electrons rotating about the inner group and which are much less rigidly attached to the atom. To explain why the physical and chemical properties of the atom do not go hand in hand with X-ray and γ -ray phenomena, it is further assumed that (a) the outer group of electrons are responsible for the physical and chemical properties of the atom; that (b) the inner group determines the phenomena associated with radioactivity; and that (c) the positive nucleus mainly determines the mass of the atom, while the groups of negative electrons are but a small multiple of the atomic weight. Hence, the inner group can be called *fixed electrons*; and the outer group, *valency or mobile electrons*. Mosely's atomic number is supposed to represent the number of positive charges or free *protons* which are concentrated at the centre of the atom. Each atom being neutral has the same number of negative charges or *electrons*. Thus the oxygen atom has 8 of each; nitrogen, 7; carbon, 6; etc. N. Bohr (1913) has applied the quantum theory of radiation (p. 700) to Rutherford's model atom to calculate the size of the atom, and to explain the series of lines found in the spectra of some elements. The results deduced from the mathematical theory based on Rutherford's atom have been very encouraging.

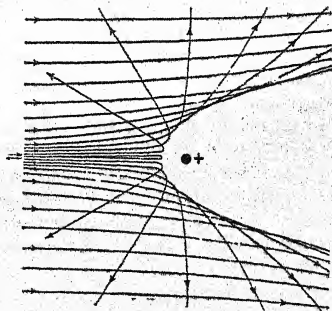


FIG. 355.—Paths of α -particles in the vicinity of an Atom Nucleus.

J. W. Nicholson (1914) has extended Bohr's work. He calculated the possible spectral lines of nebulium, and found that his theory did not account for two of the lines which had been recorded for this element in the spectra of the nebulae. Wolf of Heidelberg later found that these two lines in the spectrum of the ring nebula in Lyra had a different origin from the true nebulium lines; Nicholson also

predicted the existence of a new nebular line not previously noticed, which Wright of the Lick Observatory found on photographic plates taken years previously, but which was so faint that it escaped notice.

The periodic law.—Ever since J. B. Dumas (1851) wrote :

Every chemical compound forms a complete whole. Its chemical nature depends primarily on the arrangement and number of the constituent atoms, and to a less degree on their chemical nature,

it has been considered a fundamental principle in chemistry that substances similarly constituted have similar properties. If stable aggregates of electrons are similarly constituted, their properties, within certain limitations, will be similar even though the atomic weights be different. In W. Kossel's modification of J. J. Thomson's ring system, the two- and eight-ring systems are assumed to be the most stable, and while the outer ring of electrons determines the valency of the atom, the inner stable rings determine the characteristic properties of the element. The first series in the periodic table has an inner ring with two electrons and an outer

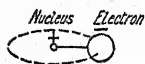


FIG. 356.—Neutral Hydrogen Atom (Diagrammatic).

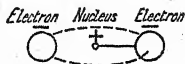


FIG. 357.—Negatively Charged Hydrogen Atom.

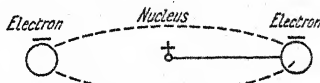


FIG. 358.—Neutral Helium Atom.

ring with from zero to seven electrons, the addition of one more electron to the outer ring makes a stable ring from which no more electrons can escape :

First series . (2) (2)1 (2)2 (2)3 (2)4 (2)5 (2)6 (2)7
 Second series . (2,8) (2,8)1 (2,8)2 (2,8)3 (2,8)4 (2,8)5 (2,8)6 (2,8)7
 Third series . (2,8,8) (2,8,8)1 (2,8,8)2 (2,8,8)3 (2,8,8)4 (2,8,8)5 (2,8,8)6 (2,8,8)7

Hence when the elements are arranged in the order of increasing mass, there is a certain similarity in the grouping at certain intervals. Thus, the properties connected with a two-ring group can only recur at intervals; similarly with more complete systems. Hence, we can divide the various groups of electrons into families such that any one family is derived from the preceding members by the addition of another ring of eight electrons. This idea gives a rather definite conception of the meaning of the periodic law. A periodic law thus appears as a necessary consequence of the hypothesis that atoms are built of stable systems of concentric rings of electrons; for obviously, *certain rings of electrons recur periodically with an increase in the number of electrons which make up the atom, and atoms*



FIG. 359.—Positively Charged Helium Atom.

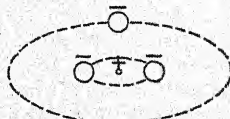


FIG. 360.—Neutral Lithium Atom.

with related rings must possess many common properties in virtue of the similarity in the grouping of some of the rings. Thus, J. J. Thomson has demonstrated that the spectra of such groups would be in many ways similar.

The calculations and the observations on the spectra of the elements

under different conditions show that the neutral hydrogen atom is probably constituted with one electron as depicted in Fig. 356, while the negatively charged atom has another electron, Fig. 357. The neutral helium atom has 2 electrons, and a doubly charged positive nucleus, Fig. 358; and the positively charged helium atom has one electron, Fig. 359. The neutral lithium atom is compounded of 3 electrons, with a triply charged positive nucleus, Fig. 360. N. Bohr's arrangement of the electrons in the atoms of elements of low at. wt.—from hydrogen to chromium—is shown in Table LXVI. The iron-group occurs at the end of the third period of eight

TABLE LXVI.—N. BOHR'S ARRANGEMENT OF ELECTRONS IN THE ATOMS OF ELEMENTS OF LOW ATOMIC WEIGHT.

Atomic number.	Element.	1st ring.	2nd ring.	3rd ring.	4th ring.	5th ring.
1	Hydrogen	1	—	—	—	—
2	Helium	2	—	—	—	—
3	Lithium	2	1	—	—	—
4	Beryllium	2	2	—	—	—
5	Boron	2	3	—	—	—
6	Carbon	2	4	—	—	—
7	Nitrogen	4	3	—	—	—
8	Oxygen	4	2	2	—	—
9	Fluorine	4	4	1	—	—
10	Neon	8	2	—	—	—
11	Sodium	8	2	1	—	—
12	Magnesium	8	2	2	—	—
13	Aluminium	8	2	3	—	—
14	Silicon	8	2	4	—	—
15	Phosphorus	8	4	3	—	—
16	Sulphur	8	4	2	2	—
17	Chlorine	8	4	4	1	—
18	Argon	8	8	2	—	—
19	Potassium	8	8	2	1	—
20	Calcium	8	8	2	2	—
21	Scandium	8	8	2	3	—
22	Titanium	8	8	2	4	—
23	Vanadium	8	8	4	3	—
24	Chromium	8	8	4	2	2

elements. Here the elements of neighbouring groups show similar chemical properties, indicating that the configurations of the electrons in the elements of this group differ only in the arrangement of the inner electrons. After the iron group, the period is no longer 8 but 18, suggesting that the elements of higher at. wt. have a recurrent configuration of 18 electrons in the innermost ring. Such a ring would probably be unstable, and the electrons may be arranged in two parallel rings. The presence of the rare earth group indicates that an alteration of the innermost rings occurs. In the case of multivalent elements, it is possible that electrons can move from the outer to an inner ring, so that multivalency corresponds with a kind of infra-atomic tautomerism. Thus, with univalent copper, the arrangement for the atoms could be 2, 8, 8, 10, 1, while with bivalent copper

the arrangement could be 2, 8, 8, 9, 2. This recalls A. Grünwald's definition of an atom (1887):

A chemical atom is a complex of exceedingly many movable particles, which are elastic, but so intimately connected together that no chemical process which comes under our consideration is capable of severing this union and breaking the atom into fragments. The parts of the atom are not conceived of as absolutely immutable any more than the atom itself, but as capable within finite limits of undergoing modifications, which have definite relations to their mutual reactions.

It will be observed that one oxygen atom is equivalent to the substance of 16 hydrogen atoms, for there are 8 free protons (+ charges) in the nucleus, and 8 electrons in the space outside, as well as either 8 hydrogen atoms or 2 helium atoms packed into the nucleus. It requires the substance of 4 hydrogen atoms to make 1 atom of helium, for there are 2 free protons in the nucleus, 2 electrons outside, and 2 more hydrogen atoms in the packing. The comparative weights of these atoms are oxygen 16, helium 4, and hydrogen not 1, but 1.008. The reason why the relative weight of helium is not 4.032, and of oxygen, not 16.128, is explained by assuming that when hydrogen is consolidated into helium, the fractional mass—about eight one-thousandths of each hydrogen atom—is transformed from matter into energy.

The misfits in Mendeléeff's table.—It is even possible to see how elements with similar properties may fall into the "wrong" group in Mendeléeff's table if classed solely by their atomic weights. Starting from sodium in the scheme on p. 992, we can see that if the elements are really formed by the aggregation or condensation of electrons (p. 1014), the sodium atom might collect more electrons until, say, two different stable systems capable of separate existence are formed. Thus, sodium might furnish two subgroups—potassium and copper. The properties of an element are supposed to be determined by the structure of rings of electrons, and differences in the properties of members of the subgroups are due to the differences in the internal structure of the atoms although the atoms probably possess some rings in common. Each subgroup, by a further condensation of electrons, forms the succeeding family members indicated in the vertical columns of the table. Consequently, it is quite true, in a general way, to say that the elements were evolved in the order of their atomic weights, but they must also have evolved in groups down the vertical as well as along the horizontal lines. Hence, as J. N. Lockyer observed (p. 1007), the elements do not always appear in the cooling stars in the order of their atomic weights. In virtue of this multiple growth hypothesis of A. C. and A. E. Jessup (1908), it is to be expected that in some cases an element in any particular group may contain more or less electrons, and hence have a slightly greater or less atomic weight than adjacent elements in the next succeeding group.

Working from an atom of the Rutherford type, J. W. Nicholson (1914) has calculated spectral lines for vibrating electrons in systems with an assumed atomic number 1e, 2e, 3e, 4e, 5e, and 6e, corresponding with Moseley's atomic number, and with the respective atomic weights:

Element	Phr	—	—	Nu	Pf	Arc
Atomic number	1e	2e	3e	4e	5e	6e
Atomic weight	0.082	0.327	0.736	1.31	2.1	2.9

The lines corresponding with 1e exist in the spectra of nebulae and have been assumed to belong to an unknown primitive element called *proto-hydrogen*, Phr;

the lines corresponding with $2e$ are present in nebulae and in Nova Persei; similarly faint lines corresponding with those computed for $3e$ are found in nebulae; $4e$ belongs to *nebulium*, Nu ; the lines for $5e$ have not been found in nebulae, but those in the solar corona have been supposed to represent a non-terrestrial *proto-fluorine*, Pf ; the lines for $6e$ exist in the spectra of nebulae and have been supposed to represent an unknown element *arconium*; Arc lines corresponding with $7e$ are also found in nebulae.

Electronegative and electropositive characters.—The majority of those working on the structure of the atom believe that the outer ring of the electrons determines the periodicity and valency of the atoms of the elements. C. A. Kraus argues as follows: The outer electrons are held loosely and are able to move from atom to atom. These electrons are very sensitive to changes in condition, such as temp., press., the presence of other atoms, etc. So weak is the bond uniting the electron to an atom, that more electronegative atoms may remove it from the original atom in question. *The less tendency the metal has to retain its electron, the more electropositive it becomes, and the more readily does it in general react chemically.* Ordinarily, the positive and negative constituents of a compound are held together through the medium of the electron. Under certain conditions, however, the electrostatic force acting between the metallic atom and its electron becomes weakened to such an extent that the negative constituent escapes, carrying the electron with it. The same result may be obtained at high temp. with the fused salt or even with the solid compound. That is when sodium and chlorine combine, the sodium atom gives up an electron to the atom of chlorine, and the atoms are thus held together by the electrostatic forces between the positively charged residue of the sodium atom and the negatively charged atom of chlorine. In a soln. of high dielectric constant such as water, these electrostatic forces are weakened to such an extent that ionization occurs resulting in the formation of Na^+ and Cl^- ions. Naturally the properties of those ions are radically different from those of metallic Na and gaseous Cl_2 as we know them. Consider the properties of the systems with a 2-electron ring on the inside, supposed, in W. Kossel's system, to be

Li	Be	B	C	N	O	F
(2)1	(2)2	(2)3	(2)4	(2)5	(2)6	(2)7

When the first member with three electrons is subjected to a small disturbance, an electron is detached from the outer ring. When one electron is lost, the residue has a positive charge, and behaves like a positively charged ion, that is, like an electropositive or basic element. The system with four electrons is more stable than the one with three, and it will not be so readily broken as the preceding. This means that this system will not be so basic as the preceding. Similarly the system with five electrons will be less basic than the one with four; and six will be less basic again than five. In the six-electron system, the stability is so great that there is little danger of losing electrons from the outer ring, and an electron could lie on the surface of the system without breaking a ring. In that case, the system would receive a negative charge and behave like an electronegative or acidic element. Hence, the electronegative elements may be regarded as possessing neutral atoms with a tendency to absorb negative electrons from without, and electropositive elements are regarded as neutral atoms with a tendency to part with negative electrons. The electronegative

character increases with increasing atomic weight until the system with 9 electrons is reached. This corresponds with the fact that in Mendeléeff's list of elements the electronegative property is practically zero at the end with the alkali metals, and gradually becomes more marked with increasing atomic weight until, at the halogen end, it attains its maximum value.

Null-valency.—As just indicated in W. Kossel's ring system, modified from that of J. J. Thomson, it is assumed that the two- and eight-ring systems are the most stable. Such rings are not readily broken down. If an electron escaped from such a system the positively charged residue will attract surrounding electrons so that one will immediately dart back to it, and reform the original system electrons. The system so formed will break up as before and the same cycle of changes would be repeated over and over again. Hence the system will not remain permanently charged; directly it loses an electron another takes its place. Such an atom would be unable to retain a positive or negative charge permanently; and it would not be able to enter into chemical combination. Consequently, it would behave like the members of the group of inert gases.

Valency.—The system with ten electrons resembles the one with three in respect to stability, etc., and the system with eleven resembles the one with four, and so on. The system of three electrons, with two electrons in the inner ring, will be most electropositive of the series. It can lose but one electron because, if it lost two, the residual one-electron system would be unstable, and would immediately attract an electron. Hence, the three-electron system can lose one electron and form a system carrying unit charge of positive electricity. This means that such a system is univalent. The four-electron system would not be so ready to part with electrons as the three-electron system, but it could afford to lose two electrons since it is not reduced to the unstable one-electron group until it has lost three electrons. The group with four electrons must therefore be bivalent. Against this, however, it must be remembered that positive ray analysis shows that there is no relation between the number of electrons an atom can lose and the valency of the element; thus the inert gas helium readily loses two electrons; the argon atom can carry one, two, or three positive charges; krypton, four; etc.

Chemical combination.—Suppose that we have an atom of the type (2,8)1 and one of the type (2,8)7; the former would be illustrated by an atom of sodium, and the latter by an atom of chlorine; the former can *lose* one negative electron, and the latter can *gain* one, but not more than one. The electrons which escape from the sodium atom can find a home on the chlorine atom, and if an equivalent number of both be present, each of the sodium atoms will acquire a positive charge, and each of the chlorine atoms a negative charge. The oppositely electrified atoms will attract one another, and form a compound NaCl. When electropositive and electronegative atoms are brought into proximity, the former loses and the latter gains an electron, as a result, the former acquires a positive and the latter a negative charge; and the force of chemical affinity is the attraction exerted by the electropositive atom for the electron it has lost to the electronegative atom. Similarly, if equivalent quantities of atoms of the type (2,8)2 and (2,8)7 had been mixed, and this would be illustrated by calcium and chlorine atoms, the atoms of calcium would each lose two negative charges, and each of the chlorine atoms would gain

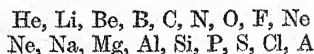
one negative charge. Thus a neutral system would be formed by the combination of two chlorine atoms with one calcium atom and CaCl_2 would result. Hence, says Thomson, from this point of view a univalent electropositive atom is one which can lose one and only one electron to form a stable system under conditions which prevail when chemical combination is taking place. Similarly, *mutatis mutandis*, with univalent electronegative atoms, bivalent atoms, etc. Thus, the affinity of an atom depends upon the ease with which an electron can escape from or be received by the atom. This may be influenced by the conditions prevailing at the time chemical combination is taking place. If the atoms be diffused in a good conducting medium it would be easier for a liberated corpuscle to resist being pulled back to the original atom than if the atoms were diffused in a non-conducting medium. Hence, the valency of an atom may be influenced by the physical conditions under which it is placed. The ability of an atom to enter into chemical combination depends upon its power of acquiring a charge of electricity. Thus chemical affinity is electrical affinity, or, as H. Davy expressed it in 1819, "electrical and chemical attractions are produced by the same cause acting in one case on particles, and in the other on masses."

This can be represented another way. If an electron on the outside of an atom is brought into contact with another atom in which is present an orbit requiring less energy to maintain the rotation of the electron, it will jump from the one to the other—that is, from the one requiring more to the one requiring less energy—so as to maintain the most permanent available condition. Each rotating electron, so to speak, endeavours to find the easiest job within its reach—the orbit which requires the least energy to maintain its motion. These jumps of the outside electrons from one atom or group of atoms to another constitute the whole of chemical phenomena. The jumps do not always occur from the higher to the lower potential, because energy may be supplied as heat or electricity, or in other guises. When an electron passes from one orbit to another within the atom the difference in energy required to maintain it in the first orbit over that required to maintain it in the second is given off. This difference is always the same; it is the quantum of energy, ϵ , of p. 700.

Abegg's normal and contra valencies.—We have just seen that the (2,8) system can acquire one and only one positive charge by the loss of one negative electron; but it is conceivable that additional negative electrons could be forced into the system so that the total number of electrons increase to (2,8)2, (2,8)3, . . . , (2,8)7; and at the same time the systems would become more and more stable. If an additional electron were forced into the (2,8)7-system, an unstable system with (2,8)8 electrons would be formed. Consequently, (2,8)7 is the greatest number of negative electrons which we can hope to force into the (2,8)-system to furnish a stable system with an electronegative valency of 7—the electropositive valency of the (2,8)1-system is one. Similarly, *mutatis mutandis*, with the other systems, and Thomson has tabulated the properties of the systems containing (2,8) to (2,8)7 electrons as follows:

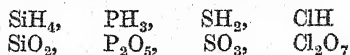
Number of electrons	(2,8)	(2,8)1	(2,8)2	(2,8)3	(2,8)4	(2,8)5	(2,8)6	(2,8)7	(2,8)8	
Valency	+0	+1	+2	+3	+4	+3	+2	+1	+0	
	-8	-7	-6	-5	-4	+5	+6	+7	+8	
	Electropositive.				Electronegative.					

This sequence of properties imitates that observed with the elements:



The first and last members of the series only are null-valent; the second set is univalent electropositive, and the last but one univalent electronegative; the third is bivalent electropositive, and the last but two bivalent electronegative, etc.

This recalls an observation of D. I. Mendeléeff to the effect that the sum of the maximum oxygen and hydrogen valencies of the elements in certain groups of the periodic series is equal to 8; *e.g.*,



In ignorance of the number of electrons in successive rings, it is inferred that the number of mobile electrons in the outer ring cannot exceed eight, and this represents the greatest charge which has yet been observed, the minimum is zero. It is further assumed that there is a tendency for these electrons in every atom to assume either the maximum or the minimum value. Thus, if an atom has five mobile electrons, it can either lose all five and acquire a positive valency of five, or it can attract three more electrons from without, make its complement up to eight, and thus acquire a negative valency of three. Which of these two possible effects will take place, will depend upon the nature of the neighbouring atom. Thus, if a phosphorus atom be surrounded by hydrogen atoms, under the right conditions of temperature, etc., its behaviour would correspond with its gaining three electrons, becoming electronegative, and forming PH_3 with a maximum valency of three; on the other hand, if it be in contact with a strongly electronegative element like chlorine, it appears to lose five electrons, become electropositive, and to form PCl_3 or PCl_5 with a maximum valency of five. Variable valency here meets a rational interpretation, for the whole of the available electrons need not be always removed together. The work done in removing a second electron must be greater than is involved in removing the first. So with phosphorus, the force exerted by the electronegative chlorine atom may suffice to drag three electrons from the phosphorus atom, but only when the conditions are favourable is it possible to drag off another two. This does not explain the rarity of intermediate valencies between three and five.

From purely chemical considerations, too, R. Abegg (1902-4) was led to assume that every element possesses a maximum valency of 8 made up of positive and negative components according as the element is acting as an electropositive or electronegative constituent of a compound. The two valencies of opposite polarity are called normal valencies and contra valencies. The normal valencies are supposed to be the stronger, and correspond with the usually accepted maximum valencies of the elements. In the case of the metals the normal valencies are positive and the contra valencies negative; while in the case of the non-metals, the normal valencies are negative, and the contra valencies positive. Thus:

	Na	Mg	Al	Si	P	S	Cl
Normal valencies	+1	+2	+3	+4	-3	-2	-1
Contra valencies	-7	-6	-5	-4	+5	+6	+7

Thus, chlorine is univalent in HCl where it is combined with electropositive hydrogen; but it has its maximum septavalency when united with electronegative oxygen in chlorine heptoxide. A. Werner has also shown how elements have a different valency according as they are united with electropositive or electronegative elements. The relative strengths of the two kinds of valency depend upon the nature of the associated atoms. The alkali metals are so strongly electropositive that they show little if any sign of an electronegative valency; and fluorine appears to be too strongly electronegative to show a positive valency, for it forms no compound with oxygen. When the conditions are such that the latent contra valencies become operative, complex "molecular" compounds may be formed.

The fact that in a given family the elements with the greatest atomic volume are usually most electropositive and least electronegative, is taken to mean that when two atoms of different sizes are in contact, the attraction of the smaller atom for the electrons in the larger atom is greater than the attraction of the larger atom for the electrons in the smaller, so that the resultant force will drag corpuscles from the larger to the smaller atom. Again, since the atomic volume increases with the atomic weight, the electropositive character of the elements in a given family increases with the atomic weight. Examples have been indicated in reviewing the different family groups of elements.

Radioactivity.—Assuming that the atoms are built of systems of rotating rings of electrons, the configuration of a system of rotating rings of electrons is dependent not only upon the number, but also upon the energy of the speed of rotation of the rings. Four electrons, for instance, may arrange themselves at the corners of a square or a tetrahedron. If a four-square system of electrons be rotating faster than a certain critical value, they will be stable; but if their velocity falls below the critical value, the arrangement will become unstable, and the electrons will suddenly arrange themselves in the form of a tetrahedron. Similarly, if a spinning top be rotating faster than its critical value, it will remain stable in a vertical position, but if the speed of rotation falls below this value, the top becomes unstable, and falls down; in doing so, it gives up a considerable amount of energy. These analogies can be extended to complex groups of electrons, say a radium atom. Owing to the radiation of energy, the kinetic energy of the electrons is gradually reduced, and the velocity of the spinning rings of electrons must be slowly diminishing. When the velocity approaches the critical value, the configuration of the system may be modified, and this is accompanied by an increase in the rate at which kinetic energy is lost by the radiation. When the velocity reaches the critical value, the configuration becomes unstable, an atomic cataclysm, or an atomic explosion occurs, and a number of electrons are detached from the original assemblage. In other words, the atom disintegrates and a part of the atom is shot off to form two or more groups of electrons. This corresponds with the emission of α -rays and emanation from radium. As G. H. Darwin (1905) expressed it, the laws of electricity in motion proves that a community of electrons, as pictured by Thomson, must be radiating or losing energy, and therefore the time must come when it will be run down as a clock does. The aggregate will then spontaneously change into another system (or element) which needs less energy than was required in the

former state. If the atoms of the emanation are built of electrons of the same type as the original atom, the process may be repeated with the sub-atoms, and so produce a series of degradation products with a long or a short life. This shows that with atoms of a special kind—*e.g.* high atomic weights—the gradual reduction of the kinetic energy of the electronic motion might produce instability within the atom. The rate of decrease of the kinetic energy may take thousands of years before it reaches its critical value, or it may take place in a very short time. The atoms of the several elements thus represent different aggregates of electrons which have proved by their stability to be successful in the struggle of the elements for a separate existence. This recalls the saying of an old Greek writer: Forms unsuited to their conditions must perish, and forms suited to their conditions are maintained.

An enthusiastic writer declares that "the corpuscular theory of Thomson has enabled chemists to explain the periodic law, valency, chemical action, etc., in terms of known facts." We are also told that more is known about the atom than about matter *en masse*. These are surely based on superficial views which are the result of confusing fact and fancy. The electronic model atom has but a remote analogy with the real atom. Analogy may be an invaluable aid to description, but it cannot *prove* a single fact. Thomson quite recognized the inadequacy of his model atom, for he points out that the number of electrons corresponding with a particular property will doubtless be different if the electrons are distributed in three-dimensional space instead of in concentric rings. With all its imperfections, this hypothesis throws a most interesting light on Mendeléeff's series and on the possible nature of an atom; though matter still remains an inscrutable and impenetrable mystery. We are still puzzled, for, as W. Crookes (1886) has said: "The list of elements extends before us as stretched the wide Atlantic before the gaze of Columbus, mocking, taunting, and murmuring strange riddles which no man has yet been able to solve." Each forward step in our knowledge carries us backwards to a vaguer and remoter past.

§ 14. Electronic Hypotheses of Chemical Combination and Valency.

Truth is the daughter of time.—FRANCIS BACON.

In recent years there has been a constantly growing tendency to apply the electronic theory to explain chemical phenomena. It is assumed that electricity has an atomic structure (p. 360), and is not continuous; and that positive and negative units of electricity are associated as constituent parts of the chemical atom. Consequently, the association of atoms to form molecules involves an investigation of the structure of fields of force exerted on the outer surface of the atom. The different hypotheses usually take one of two forms:

1. Every chemical bond between two atoms involves a *transfer* of one negative electron from one atom to the other, so that the one atom is charged positively and the other negatively—*vide* p. 361.

2. If the electrons are actually displaced, J. Stark (1908) considers that some phenomenon resembling the selective absorption of light takes place, and that when the electrons return to the atom, a phenomenon resembling fluorescent radiation occurs.

The atom can be regarded as constituted of a positive nucleus, with n electrons rotating in orbits about a central nucleus—*dynamical hypotheses*—e.g. W. Kossel, N. Bohr, etc.; or, the electrons may be confined to particular parts of the structure—*statical hypotheses*—e.g. G. N. Lewis, I. Langmuir, etc. The former are generally considered to be nearer the truth and are specially favoured by physicists. J. Larmor, for example, compared the atom with a clock, the outer electronic shell being the pendulum, maintained in continuous vibration by the energy of the core imparted through an unknown escapement mechanism which parcels out quanta of energy.

(a) In one form of Stark's hypothesis (1908) the units of positive electricity are supposed to be situated on the surface of the atoms, not as uniformly charged zones, but rather concentrated at certain points, and the neutralizing electrons are assigned definite positions on the surfaces of the atoms. The negative electrons hold the positive atoms each to each, as illustrated diagrammatically in Fig. 361, where the dotted lines represent tubes or lines of force.

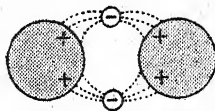


FIG. 361.

(b) In J. J. Thomson's hypothesis (1914) the electric forces which keep the atoms together originate in a *displacement* of the positive and negative electricity in each atom; and, as a result, each atom acts like an electric doublet and attracts another atom, much in the same way that two magnets attract one another.

J. J. Thomson's theory starts from the assumption that an atom is an aggregate of negative electrons with a sphere of positive electrification. The charges may or may not counterbalance one another, leaving the atom either electrically neutral or polarized, that is, charged with a surplus of positive or negative electricity. A pair of balanced charges is called an electric doublet, and W. Sutherland suggested the hypothesis that the peculiar force exerted by an atom is the same as if each atom carried an electric doublet whose moment is characteristic of the atom. Sutherland's atoms therefore characterize compounds of the first class, but not the second.

The field of electric force about the molecules of a compound.—From the known properties of electrified bodies, it is inferred that a field of electric force must envelop each atom no matter whether the opposite charges wholly or partially counterbalance one another. Even an electrically neutral atom will have a field of force which (i) will not be uniformly distributed about the atom; and which (ii) will probably diminish in intensity more rapidly than the inverse square of the distance. The field of force emanating from the charges in the atom will give rise to (a) *intramolecular attractions* whereby the charged atom affects the other atoms associated with it in the molecule; but it will also give rise to (b) *intermolecular attractions* because the field of force of an atom will also attract the atoms in other molecules, and so produce phenomena like the surface tension of liquids, latent heat of evaporation, cohesion of liquids and solids, etc. Accordingly, chemical compounds can be divided into two classes:

I. *Molecules with uncharged or neutral atoms.*—The individual atoms of the molecule are neutral, for the constituent atoms are charged with equal

amounts of positive and negative electricity, and there is no surplus or excess of the one kind over the other. *E.g.* H_2 , O_2 , He , N_2 , Cl_2 , CO , CO_2 , CS_2 , CCl_4 , C_6H_6 , N_2O , etc.

II. *Polarized molecules, or molecules with charged atoms.*—The individual atoms within the molecules carry an excess of positive or negative electricity so that the whole molecule is charged, and, accordingly, exerts larger forces upon the atoms of neighbouring molecules than would be the case if the atoms were neutral. The chemical reactivity, specific inductive capacity, surface tension, cohesion, tendency to form molecular complexes, and other qualities of such compounds are found to be highly developed. *E.g.* H_2O , NH_3 , HCy , SO_2 , HCl , CH_3OH , $\text{C}_2\text{H}_5\text{OH}$, CH_3Cl , etc.

The distribution of the electric charges within the molecule.—Each element has its own specific attraction for negative charges, thus, sodium exerts a less attraction for negative charges than chlorine; otherwise expressed, sodium is more electropositive than chlorine. With carbon monoxide, although the opposite charges on the two atoms neutralize one another, yet the oxygen has a greater attraction than carbon for a negative charge. Hence, a certain amount of work is available in the transfer of a negative charge from the atom in which the attraction is weak to an atom in which the attraction for the negative charge is strong. Accordingly, when a neutral electropositive atom is united with an electronegative atom, the negative charge tries to pass from the former to the

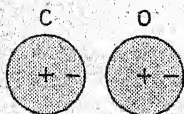


FIG. 362.—Carbon monoxide, CO , with neutral atoms.

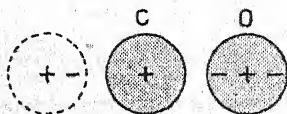


FIG. 363.—The charging of a Carbonyl, CO , radicle.

latter and electrify it negatively. Why does not the transfer always take place? Thomson assumes that electricity has an atomic structure, and must pass in definite units from one point to another, and therefore a whole charge must be transferred or none at all. He illustrates the action by diagram, Fig. 362. Each atom C and O carries a balanced positive and negative charge. The O is more electropositive than the C , and there is therefore a tendency for a negative charge to pass from the C to the O atom; but this tendency does not suffice to overcome the attraction of the positive charge carried by the C atom. This latter is diminished by bringing up another atom represented by the dotted line, Fig. 363. This helps the passage of the negative charge across to the O atom. Thus, the atoms of a compound may be uncharged when isolated, and yet become charged when the compound is in a liquid state, or when the atoms are combined to form more complex molecules. The term intramolecular ionization, by the way, is applied by J. J. Thomson to the process by which the atoms of a molecule get charged electrically.

Leyden jar analogy.—J. J. Thomson compares the two atoms of, say, carbon monoxide with the two coatings of a Leyden jar; so that the whole molecule behaves as if it were a Leyden jar of small capacity. It is known that the smaller the capacity of a jar, the greater the amount of energy required to transfer, say, 7 units of electricity into the jar. If the available energy be less than this, the transfer of, say, a single negative charge from a carbon to an oxygen atom will not

occur, and *vice versa*. By increasing the capacity of the "jar," a transfer might occur which would otherwise be impossible. Hence, if another molecule be added to a certain atomic system, its capacity might be increased and accordingly less work might be required to transfer a charge from an electropositive to an electronegative atom.

Platinum tetrachlorodiammine, $\text{PtCl}_4(\text{NH}_3)_2$, is not an electrolyte, because its atoms are electrically neutral. Add another molecule of ammonia to get $\text{PtCl}_4(\text{NH}_3)_3$, and the resulting compound is an electrolyte with a molecular conductivity of 97; this number rises to 228 with the addition of another molecule of ammonia to form $\text{PtCl}_4(\text{NH}_3)_4$; and up to 553 when two more molecules of ammonia are added to form $\text{PtCl}_4(\text{NH}_3)_6$. Thus, as the molecules of ammonia are increased, the charges carried by the atoms increase rapidly. Carbon monoxide is neutral, but when it enters as a carbonyl radicle into organic compounds—like acetone $(\text{CH}_3)_2\text{CO}$; aldehyde, CH_3COH ; etc.—it raises their specific inductive capacity because the CO radicle becomes charged on entering the compound. The proximity of other atoms increases the capacity of the system sufficient to allow the transfer in question. Mere contact with the other atoms promotes intramolecular ionization. In some cases, some of the radicles or atoms may carry one or more charges, and other atoms may be electrically neutral. The hydroxyl radicle, OH, may exist in a molecule in two states according as the oxygen atom carries one or two negative charges. If the oxygen atom carries one charge, the radicle $\text{O}^- - \text{H}^+$ will form an electric doublet; and if two charges—as is probably the case with water—there must be a positive charge on some other radicle not the hydroxyl group, and there will be two doublets in the molecule. J. J. Thomson suggests that the difference in these two states determines whether the OH radicle acts as an acid or as a base. Acidic hydroxyl is represented by $\text{O}^- - \text{H}^+$, and basic hydroxyl (of water) $\text{O}^{2-} - \text{H}^+$, for if the molecule $\text{RO}^- - \text{H}^+$ be in solution and surrounded by H^+ and $\text{O}^{2-} - \text{H}^+$ ions, arising from the dissociation of water, the H^+ ion of the given molecule would unite with the $\text{O}^{2-} - \text{H}^+$ of the water, and leave an excess of H^+ ions in solution so that the molecule $\text{RO}^- - \text{H}^+$ would have acidic properties; on the other hand, in the molecule $\text{R}^+ - \text{O}^{2-} - \text{H}^+$, the $\text{O}^{2-} - \text{H}^+$ of the molecule would unite with the H^+ from the dissociation of the water and leave an excess of $\text{O}^- - \text{H}^+$ ions in the water, and the molecule would have basic properties. The more electropositive the element R in the ROH molecule, the more likely is a corpuscle to be transferred from the R to the O atom, and accordingly the more likely is the oxygen to acquire a second charge and show basic properties.

Valency electrons or corpuscles.—J. J. Thomson (1914) further suggests that each atom may contain negative electrons related to the atom in two ways: (i) *Fixed electrons*. Some of the electrons are firmly fixed at the core of the atom; they are not free to adjust themselves so as to cause the atom to attract other atoms into its neighbourhood; and they take no part in chemical reactions. (ii) *Mobile or valency electrons*. Some of the negative electrons are supposed to be located near the surface of the atom, and they are free to move about and set themselves into position under the influence of external electrical fields. They are linked with the central positive charge of the atom by lines of force. The mobile electrons, which J. Stark (1908) called *Valenzelektronen*, enables the atoms to hold on to one another and form a bond between joined atoms. The number

of mobile corpuscles determines the valency of the atom so that a univalent atom has one, a bivalent atom two, and a septivalent atom has seven mobile electrons. The mobility of the electrons is an essential condition necessary for one atom to exert any considerable attraction on another. For example, if an atom with one mobile negative electron—Fig. 364—and another atom with its negative electron—Fig. 365—were in contact,

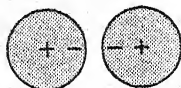


FIG. 364.—Repulsion.

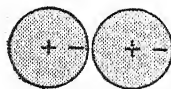


FIG. 365.—Attraction.

the two atoms would attract or repel one another as illustrated in the diagrams. If, however, the electrons are free to move, they will be driven apart until the positive charge of the one atom is nearer to the negative charge of the other, and the atoms would then attract one another. The negative corpuscle has very little inertia and, if free, can be driven round at once. The initial repulsion would therefore be momentary and attraction would be the final result. If the corpuscle was fixed, it could not swing round without carrying the whole atom, which is a comparatively heavy body difficult to move. As a result, when the corpuscles are mobile, the atoms are kept together, but when the electrons are fixed, the atoms will have comparatively little attraction for one another. When the negative corpuscles of the atom are so firmly fixed as to be incapable of exerting any great attraction on other atoms surrounding it unless they are in particular positions, the valency of the atom is said to be saturated, for the satisfaction of a valency is essentially the fixing of one of the mobile negative particles.

Fixing the valency electrons.—Assume that each negative electron in an atom is the origin of a line or tube of force. If the atom is alone, the line of force will return to the same atom and end on its own positive charge as illustrated diagrammatically, Fig. 366. The negative electron is still free to move. Consequently, an electron will not be fixed when the atom is by itself, but only when the line of force from the corpuscle is anchored to an opposite charge on another atom *B*, Fig. 368. The electron is then deprived of its mobility, and is unable to attract another atom.

Case I. If the two atoms be electrically neutral, a similar tube of force will pass from each negative electron on each atom to the positive charge on the other atom, as illustrated diagrammatically (Fig. 368). Hence, with neutral atoms, for each tube of force which leaves an atom there will be a return tube—as many will go out as come in. If the atoms are each bivalent, two tubes of force will leave each atom and two return; in general in compounds with neutral atoms the total number of lines of force between the constituent atoms is double the chemical valency. This means that J. J. Thomson



FIG. 366. — Tube of Force of a Mobile Corpuscle.

would represent each chemical valency by *two* bonds between the atoms of a molecule when the constituent atoms are electrically neutral. It does not follow that if the atom *A* sends a tube of force to another atom *B*, it must receive one from the *same* atom *B*, for the return tube

may come from *another* atom. The necessary condition is that the number of tubes of force which leave an atom must be equal to the number which return. With this understanding, certain combinations can be explained which appear to be anomalous in the light of the old valency hypothesis. If hydrogen be univalent, a combination like H_3 was considered impossible on the older hypothesis. But the hydrogen molecule has neutral atoms, with two tubes of force issuing from each. The molecule H_3 , that is—



is thus quite compatible with the hypothesis, so also is a molecule H_n . Again, with monatomic silver, $AgCl$, $AgCl_2$, $AgCl_3$, and $AgCl_n$ are all possible valency compounds in this extension of the older valency hypothesis. The trouble with the hypothesis is that it apparently explains too much.

Case II. If the two atoms be charged electrically.—We have seen that an electron will be fixed when its tube of force passes out of its atom and ends upon some other atom, Fig. 368. If one of the negative electrons (Fig. 367) be dragged from one atom, A, to the atom B, the electron would follow its tube of force which left the atom A, and the tube of force would shrink up in the atom B, and only one tube of force would connect the two atoms. When a number of electrons are transferred from one atom to another, the number of tubes of force between the two atoms will be halved, because

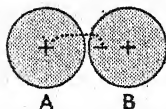


FIG. 367.—Union of Charged Atoms.

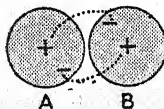


FIG. 368.—Union of Neutral Atoms.

after the transfer no tube of force will leave the electropositive atom, and none will enter the electronegative atom. The one atom, B, would thus be positively charged and the other, A, negatively charged. Hence, in compounds with charged atoms, the number of tubes of force between the constituent atoms will be equal to the chemical valency. The tube of force is then equivalent to the bond or bar of the regular structural formulæ (p. 87) employed by chemists, and the valency bond is only applicable when the constituent atoms are charged electrically. If the constituent atoms are electrically neutral each valency bond must be doubled if it is to represent a tube of force. The chemist's hyphen or bond is thus related with the physicist's tube of force.

Unsaturated compounds.—It follows from the preceding discussion that when each of the mobile electrons in an atom is anchored by a tube of force to a neighbouring atom, the average attraction of that atom for other atoms is reduced to a minimum. J. J. Thomson further shows that if the molecule of a compound is to exist in a stable form, (i) The molecules must not exert a sufficiently large attraction on the neighbouring atoms to cause them to unite and form other molecular systems. This will be the case when the work required to separate the two molecules under consideration is small compared with the average kinetic

energy of the molecule at the temperature T . (ii) The attractions between the atoms of the molecules must be great enough to prevent a separation when the molecule collides with other molecules. This will obtain if the work required to separate the atoms is large compared with the average kinetic energy of the molecule at the temperature T . These conditions are greatly influenced by temperature. Thus, with iodine vapour, the monatomic molecule may exist in a stable form at a high temperature, and pass into more complex diatomic molecules at lower temperatures.

If carbon monoxide is a stable compound when electronegative carbon is united with the more electronegative element oxygen, why does not CCl_2 also form a stable compound? Thomson answers this question as follows: As already indicated, an atom, even when saturated, must produce an electric field in its neighbourhood. This field must restrict the freedom of motion of molecules of the corpuscles in neighbouring atoms. If the electric field due to the oxygen atom in carbon monoxide be very strong, then, even though but two of the four mobile corpuscles of the carbon atom are bound to the oxygen by tubes of force, the other two will be exposed to so strong an electric field that their mobility is reduced. As a result, the attraction of the carbon atom on other atoms is reduced. The amount of residual attraction, so to speak, depends upon the strength of the electric field in the carbon atom produced by the oxygen atom. If the electric field with the CCl_2 -system be smaller than that produced in the CO -system, it follows that the latter will behave more like a saturated molecule than the former, and the carbon monoxide molecule, in consequence, might exist in a free state when the CCl_2 could not.

It might be argued that in a compound like ethane, C_2H_6 , with electro-positive hydrogen atoms, and electronegative carbon atoms, the two carbon

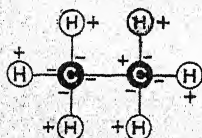


FIG. 369.—Ethane,
 C_2H_6 .

atoms are in a different electrical state, because the tubes of force from three hydrogen atoms end on each carbon atom, and a tube of force from one carbon atom must end on the other; accordingly, the potentials of the charges on the two carbon atoms must be different. This is illustrated diagrammatically in Fig. 369. Similar remarks apply to acetylene and probably also to ethylene linkages. No phenomenon has yet been noticed which would justify this conclusion in the case of ethane, but in the case of doubly and triply linked combinations, several properties—molecular volumes, indices of refraction, etc.—are markedly different from what they are with the carbon atoms all singly linked together.

This discussion has not proved these speculations, it only makes them appear plausible. The work outlined in this chapter illustrates the invaluable aid which chemistry and meta-chemistry are receiving from the disciplined imagination indicated in our preface. Naturally many have fallen before the temptation to confuse imaginary phenomena with demonstrated fact. Premature generalizations are rife, since, as a rule, the less the number of facts, the easier it is to generalize—and the more likely are those generalizations to be wrong.

The honesty of science.—Here, then, we are confronted with phantasmæ which would be banished at once if we were convinced that they

were sterile conjectures and not pregnant hypotheses. The speculations probably make the best guess yet made about the ultimate constitution of matter. The relations between hypotheses and fact, though doubtful, are not altogether contradictory. Consequently, the defective hypotheses will be persistently attacked by hostile forces until they are either abandoned in favour of more successful rivals, or developed and strengthened into a more consistent and lasting form.

This struggle for existence is the life and strength of scientific hypotheses. Science may appear to lose influence when the fallacy of a prevailing hypothesis is demonstrated; but it holds a treasured reputation for honesty of purpose by frankly acknowledging and registering its mistakes. In the words of A. Sidgwick, free from the fear of being found an impostor, science is able to challenge—and to court—correction. “As for the truth, it endureth, and is always strong; it liveth and conquereth for evermore. Truth is the strength, the kingdom, the power, and the majesty of all ages.”—1 Esdras iv.

EPILOGUE

THERE are not far from 200,000 different compounds known to chemists, and hundreds of new compounds are discovered every year. The specific properties of all these substances are described in dictionaries of chemistry, and in memoirs of the various scientific societies. The student of chemistry is not expected to be acquainted with more than a small fraction of these compounds. If a chemist discovers what he believes to be a new compound, it is possible to find if it has been previously prepared by consulting the literature just mentioned.

Herbert Spencer has properly said that in so far as the production of new compounds is carried on merely for the sake of obtaining new compounds, chemistry is not a science, but an art. The best chemist is not necessarily he who is familiar with the greatest number of compounds. Chemistry is something more than a compilation of empirical facts. Dictionaries of chemistry, not the memory, are the natural storehouses of isolated facts. The intellect is perfected not by knowledge but by exercise. The time needed for memorizing a vast medley of facts can be far more profitably spent in training the brain to think clearly and logically, and the hands to do their work skilfully and accurately. A student trusts his teacher to equip him with these essentials; and the conscientious teacher has therefore grave responsibilities. At the same time, the work of the teacher may be reinforced or hampered by an examination syllabus which the student expects to traverse.

The experience of thousands of teachers crystallized in hundreds of text-books, syllabuses of examinations, etc., is supposed to have taught teachers what facts and principles the student of general chemistry should know. But the teacher has failed in his work if he has not whet the student's appetite for more. The subsequent progress of the student in general chemistry is, however, largely determined by his intended profession, and I question if hereafter he can do better than follow the advice of "Sherlock Holmes":

I consider that a man's brain originally is like a little empty attic, and you have to stock it with such furniture as you choose. A fool takes in all the lumber of every sort that he comes across, so that the knowledge which might be useful to him gets crowded out, or at best is jumbled up with a lot of other things so that he has a difficulty in laying his hands upon it. Now the skilful workman is very careful indeed as to what he takes into his brain-attic. He will have nothing but the tools which may help him in doing his work, but of these he has a large assortment, and all in the most perfect order. It is a mistake to think that this little room has elastic walls and can distend to any extent. Depend upon it there comes a time when for every addition of knowledge you forget something that you knew before. It is of the highest importance, therefore, not to have useless facts elbowing out the useful ones.

MISCELLANEOUS QUESTIONS

1. What is meant by (a) chemical action, and (b) physical change? Is chemical action always accompanied by physical change? In the following list of changes, which are chemical? and which of a physical nature?—Souring of milk; digestion of food; liquefaction of air; freezing and evaporation of water; grinding of grain; fading of coloured fabrics; withering of leaves; weathering of rocks; melting of silver; solution of sugar in water?—*Princeton Univ. (U.S.A.)*.

2. Discuss the following quotation: "The volume of a molecule of a compound body in the gaseous state is exactly double the atom of hydrogen."

3. What were the principal difficulties in the way of the general acceptance of the atomic theory as enunciated by Dalton? Show how these difficulties were overcome.—*Sydney Univ.*

4. Suppose the mineral chromite on analysis furnished the equivalent of 50 per cent. Cr_2O_3 , on the assumption that the atomic weight of chromium is 52.1 ($\text{O}=16$). What difference in the value of a ton of chromite will be reported by two chemists if one works on the assumption that chromium has an atomic weight of 51.6 and the other 53.5? Given a 50 per cent. ore is worth £3 5s. 0d. per ton.

5. In reply to the question: "State the laws of constant composition; give an example," a candidate up for examination answered: "Chemical combination always takes place between different masses of matter, e.g. two of hydrogen to one of oxygen." Is this answer worth any marks? Why?

6. Criticize the following extract from the paper of a candidate for examination: "Gases unite chemically by measure as well as by weight; solids and liquids unite only by weight."

7. Suppose it be argued that in the experiment depicted in Fig. 13, the moisture might come (a) from the air, or (b) from the imperfect drying of the hydrogen gas, how would you proceed to modify the experiment to test if these objections stultify the inference drawn?

8. Explain the following quotation: "A fact which will not fit into reasoned and formularized schemes stands outside scientific knowledge until the right formula is found."—*J. J. Murphy (1869)*.

9. What did Bergman mean when he said that "the qualities of bodies can never be known by reasoning *à priori*"?

10. The writer of a pamphlet published in 1847 argued that water must be a simple element because "we have the great authority of Aristotle, who states water is one of the four elements"; and "chemists are at variance as to the composition of water, for according to Davy water consists of hydrogen and oxygen in the proportion of two parts of the former to one of the latter; whilst Dumas, the celebrated French chemist, says that one part of hydrogen and eight parts of oxygen form one of water." What is your opinion about these quotations?

11. In reply to a question asking for the meaning, in words, of the equation $\text{H}_2\text{SO}_4 + \text{Zn} = \text{ZnSO}_4 + \text{H}_2$, a candidate said: "Two parts of hydrogen (H_2) in sulphuric acid (H_2SO_4) are replaced by one part of zinc (Zn) and zinc sulphate (ZnSO_4), with hydrogen gas (H_2), is produced." Criticize this answer, and state how you think it ought to have been expressed.

12. The vapour densities of three substances referred to hydrogen taken as unity were 45, 70, and 25 respectively, and the percentages by weight of a certain element contained in each were, 22.22, 42.86, and 40.0 respectively. What is the probable value of the atomic weight of the element?—*Cambridge Senr. Locals. Hint*: The molecular weight of the first element is 90, and it therefore contains 20 parts by weight of the given element; while the second contains 60, and the third 20 parts by weight of the same element. Hence, if A be the atomic weight of the given element, $uA=20$, $vA=60$, $wA=20$. Hence A is probably 20.

13. "Valency is the gaseous combining ratio of hydrogen or of other elements to the unit volume of hydrogen, but extended by indirect calculations to those elements whose gaseous volume cannot be directly measured. The atomic weight is the weight of a unit volume of the elements in the gaseous condition, compared to that of hydrogen as unity."—*J. N. Sprague (1892)*. Explain this quotation more fully with examples.

14. It is said that "if a porous vessel containing a mixture of hydrogen and oxygen in any proportion be surrounded by a vacuum, the mass of oxygen which escapes in a short time is four times that of hydrogen." Expose the fallacy of this statement. *Hint*: The effect of partial pressure, which varies according to the relative proportion of the gases present, is ignored.

15. In 1703, G. Amontons published a paper on "The expansive force of air as a measure of temperature," and showed that the increase of pressure is proportional to the rise of temperature, and also to the increase of density. (a) Is the latter statement correct? (b) If air has a density of 0.00129 (gram per c.c.) at n.t.p., what must be the pressure in order that the density may be three times as great? *Hint*: (a) Since volume v varies inversely as density D (p. 32), Clapeyron's equation (p. 105) can be written $p = RTD$, etc. Hence the pressure of a gas is proportional to the temperature and density. (b) 2 atm.

16. A flask fitted with a porous plug contains equal masses of hydrogen and carbon dioxide and is exposed to the open air until half the hydrogen has passed out. Neglecting the changes of diffusion rates due to the varying partial pressure of the gases and also the back action of the air and escaped gases, find the proportion of the two gases remaining in the flask.—*R. H. Jude and H. Gossin, Physics, London, 1899.*

17. A sealed bottle can stand an internal pressure of four atmospheres and it is filled with a gas at 15° and 760 mm. Above what temperature would the bottle be liable to burst?

18. How much zinc and sulphuric acid will be required to fill a cylindrical gas holder 2 metres high and 40 cms. diameter with hydrogen gas at n.p.t.? *Hint*: The cubic capacity of the gas holder is $\pi r^2 h$, or 251.2 litres (i.e. cubic decimetres). Hence, nearly 1100 grams of sulphuric acid and 730 grams of zinc are needed.

19. It is sometimes stated that Avogadro's rule for gases is true for hydrogen and oxygen at ordinary pressure, and hence these gases unite in the proportions 2 : 1. The researches of Amagat and others on the effects of change of temperature and pressure on the volume of a gas show that hydrogen and oxygen not only deviate from the two gas laws, but their deviations are in opposite directions. Hence, adds A. Scott (1887), "it can only be by the merest chance that at our ordinary temperatures and pressures the combining volumes of hydrogen and oxygen should be exactly 2 : 1." Discuss this question.

20. Is it possible to raise the temperature of a saturated vapour at a constant pressure?

21. "The properties of a compound," said a writer in 1887, "are the sum of those of their components. This indeed is a necessity for the atomic theory, which supposes that bodies react in terms of units whose integrity remains inviolate throughout all changes." Discuss this statement.

22. Describe and explain Soret's experiments on the molecular weight of ozone. If, in Soret's apparatus, 100 c.c. of HCl gas had been mixed with oxygen, and 18 c.c. had diffused in half an hour, what molecular weight would you assign to an acid gas of which 12 c.c. (out of 100 c.c.) diffused under the same conditions?—*Owens Coll.*

23. Explain clearly the meaning of the following translation from C. L. and A. B. Berthollet's *Éléments de l'art de la teinture*, Paris, 1804: "We may be easily satisfied of the existence of oxygen in chlorine. We have only to expose to the light of the sun a phial filled with a solution of chlorine in water. Bubbles of gas are soon evolved, and the resulting gas can be collected in a suitable receiver. This gas has all the properties of oxygen. When the bubbles cease to develop, the liquid has lost its distinctive properties, and it is now water impregnated with ordinary muriatic acid and oxygen."

24. To what constituents do acids owe their characteristic properties? Does this substance impart these properties to all its compounds? What is the theory advanced to explain the facts embraced in your answer? How do we explain the fact that some acids are stronger than others? Name some other facts which the above-mentioned theory explains.—*American Coll.*

25. Describe the preparation of potassium from potassium carbonate. What products are found when water, moist air, chloric and hydrochloric acid respectively act on potassium? Give an outline of the chemical characters of each of these products.—*London Univ.*

26. State three definite reasons for placing potassium and sodium in the same class of metals.—*Cambridge Senr. Locals.*

27. A set of silver refining vats have energy equivalent to one horse-power

applied to them at an average pressure of 3.7 volts per vat. How much silver will be deposited per hour? Given 746 watts are equivalent to one horse-power. *Hints:* Since watts = amperes \times volts, the current is 200 amperes; but (p. 357) a current which will deposit 0.001118 gram of silver per second is called an ampere, hence 200 amperes will deposit $0.001118 \times 200 \times 60 \times 60 = 805$ grams per hour.

28. P. A. Guye and G. Ter-Gazarian discovered in 1906 that the most carefully recrystallized potassium chlorate contains at least 0.022 per cent. of potassium chloride. Taking account of this correction, recalculate the atomic weight of silver from the following data of Stas:—Percentage of oxygen in potassium chlorate = 39.154; ratio of silver to potassium chloride = 100 : 69.1143 (O = 16).—*Manchester Univ.*

29. Describe and explain the changes which take place when nitrogen tetroxide, enclosed in a sealed tube, is heated. Do you know any other gases which undergo similar changes when heated? In what way would you say that electrolytes, when dissolved in water, exhibit an analogous behaviour to nitrogen tetroxide when heated?—*Staffs. County School.*

30. Describe the changes that take place when the following are heated: An aqueous solution of sulphurous acid; an aqueous solution of acid calcium carbonate; an aqueous solution of 40 per cent. nitric acid; an aqueous solution of 80 per cent. nitric acid; solid ammonium nitrate; ammonium nitrite; sulphur.—*Cape Univ.*

31. (a) What are the general methods for hastening the rate of a reaction? (b) The equation $N_2O_4 = 2NO_2$ represents a system of two gaseous substances in equilibrium. What would be the effect of increasing the pressure on the system without changing the temperature?—*Dept. of Educ., Ontario.*

32. What is the action of ozone upon potassium iodide, and how would its effect be distinguished from what is produced by nitrogen peroxide upon the same salt? Calculate the weight of potassium iodide which is equivalent to one gram-molecule of ozone (K = 39.1, I = 126.8, O = 16).—*Science and Art Dept.*

33. Write a general account of the displacement of hydrogen by metals from water, from acids, and from alkalies, stating which metals act in this way and which do not. What other kinds of action may occur?—*Owens College.*

34. Give in detail practical methods by which two of the following bodies may be prepared in a crystalline state: hydroxylamine, phosphorus oxide, sodium azoimide, nitrosulphonic acid.—*Owens College.*

35. When electric sparks are passed through ammonia gas it is nearly all decomposed; when through a mixture of nitrogen and hydrogen a trace of ammonia is formed. How can either of these reactions be made practically complete? Explain the reasons for your answer. Do you know of any other instances of the same sort of phenomenon?—*Bombay Univ.*

36. According to Curtius and Schulz, hydrazine contains 87.4 per cent. of nitrogen and 12.6 per cent. of hydrogen, and it forms two chlorides; the one contains 26.96 per cent. of nitrogen, 6.01 of hydrogen, and 67.41 of chlorine; and the other contains 40.95 per cent. of nitrogen, 7.71 of hydrogen, and 51.86 of chlorine. What is the molecular formula of hydrazine? J. Biehringer, *Einführung in die Stöchiometrie*, Braunschweig, 1900. *Hint:* Take $N = 14$, $Cl = 35.5$, and $H = 1$. The formula for hydrazine, by calculation, is H_2N_2 ; the simplest empirical formula of the first chloride is NH_2Cl , and of the second N_2H_4Cl . These numbers show that the first chloride can be regarded as a compound of NH_2 with HCl , and the second chloride a compound of $2NH_2$ with HCl ; or the first might be a compound of N_2H_4 with $2HCl$, and the second a compound of N_2H_4 with HCl , etc. The second suggestion is more probable than the first, because *hydrogen chloride usually behaves in these double compounds as if it were a monad radicle—witness $NH_2.HCl$* . Hence the formulae of the three compounds in question are respectively written: N_2H_4 , $N_2H_4 \cdot 2HCl$, $N_2H_4.HCl$. Now show that this conclusion is confirmed by the vapour density of hydrazine 1.1 (air = 1).

37. By what method is sodium nitrate prepared, and what are its principal uses? How would you obtain from sodium nitrate (a) ammonia, (b) nitric oxide, (c) hydroxylamine sulphate?—*Institute of Chem.*

38. How can the nitrides of lithium and magnesium be prepared, and what is the action of water on them? Given the fact that lithium can be obtained by the electrolysis of its fused hydroxide, suggest a continuous process for the conversion of atmospheric nitrogen into ammonium salts.—*London Univ.*

39. A current passes simultaneously through acidulated water, a solution of

copper sulphate, CuSO_4 , and molten silver chloride. What substances are produced in each cell, and how many grams of each in the time that 10 cubic centimetres of hydrogen are liberated from water ($\text{Cu} = 63$, $\text{Ag} = 108$)?—*New Zealand Univ.*

40. On the assumption that the equivalent of mercury is 99.25, calculate to two places of decimals the equivalents of oxygen, hydrogen, copper, sulphur, and chlorine from the following data: mercuric oxide contains 92.59 per cent. of mercury; cupric oxide contains 70.9 per cent. of copper; cupric chloride contains 47.3 per cent. of copper; sulphuretted hydrogen contains 94.07 per cent. of sulphur; hydrogen chloride contains 97.23 per cent. of chlorine.—*London Univ.*

41. Discuss the action of pure nitric acid and of ordinary nitric acid on the metals copper and zinc, and indicate the compounds which are formed respectively under different conditions.—*Board of Educ.*

42. Compare and contrast the physical and chemical properties of magnesium with those of calcium and zinc. To which of these two metals do you consider magnesium to be more closely related? Give reasons for your answer. Given a solution containing these three metals, how would you prepare from it pure specimens of the oxides of each?—*Oxford Univ.*

43. Calculate the weight of oxygen available for oxidation in 10 grams of potassium permanganate in sulphuric acid solution.—*Sheffield Scientific School, U.S.A.*

44. How can crystallized potassium permanganate be obtained from black oxide of manganese? Explain the reactions which ensue when an acidified solution of potassium permanganate interacts with (a) sulphurous acid, (b) oxalic acid, (c) hydrogen peroxide.—*Sheffield Univ.*

45. What weight of manganese dioxide must be decomposed by hydrogen chloride in order to obtain enough chlorine to combine completely with the hydrogen evolved by dissolving 10 grams of magnesium in dilute acid? ($\text{Mg} = 24$, $\text{Mn} = 55$, $\text{Cl} = 35.5$, $\text{H} = 1$).—*Vict. Univ., Manchester.*

46. Represent by equations the action of strong sulphuric acid on each of the following substances: (a) nitre, (b) manganese dioxide, (c) formic acid, (d) copper, (e) charcoal.

47. The properties of an element depend in a great measure on the different active valencies of the element. Illustrate this principle by reference to manganese.

48. Describe the difficulties to be overcome in determining whether N_2O_3 is formed when NO and NO_2 are mixed at ordinary temperatures. What conclusions have been drawn from the results of experiments on this problem?—*Board of Educ.*

49. Describe, with all essential practical details, the preparation either of pure crystallized sodium nitrate from sodium nitrate, or of pure potassium perchlorate from potassium chlorate. What method would you adopt to ascertain the purity of the product?—*Board of Educ.*

50. State Dulong and Petit's law and explain its value in the determination of atomic weights. If the specific heats of two metals M and N be 0.25 and 0.214, and their equivalents 12 and 9 respectively, what will be the formula of their chlorides?—*Sheffield Univ.*

51. Give the composition of "pure air" by weight and by volume. Give proofs that the air is a mechanical mixture. It is intended to prepare 112 litres of nitrogen, at S.T. and S.P., from a certain chemical compound; what compound will it be best to employ, and how many grams will be required?—*Princeton Univ., U.S.A.*

52. By what experiments and reasoning has it been shown that air is a mixture of gases? By whom and how was the active constituent of air first isolated, and the nature of combustion explained?—*Sheffield Univ.*

53. Beginning with nitrogen combined in the form of nitre, state exactly how you would prepare from it (a) nitric oxide, (b) nitrous oxide, (c) ammonia, (d) nitrogen, each in a state of purity.—*London Univ.*

54. A certain metal gave the following results: (i.) 0.5 gram gave 1.396 grams of oxide, (ii.) the chloride contained 11.2 per cent. of the metal, (iii.) the vapour density of the chlorine was found to be 40, (iv.) the specific heat at ordinary temperatures was 0.46, but increased rapidly with rise of temperature. State what you can about the atomic weight of the metal.—*Oxford Univ.*

55. Explain the terms "atomic heat" and "molecular heat." The specific heat of lead sulphate, PbSO_4 , is 0.0827; that of lead 0.0309; and of sulphur

0-163. Assuming the atomic weights of lead and sulphur to be 207.1 and 32.07 respectively, calculate the specific heat of solid oxygen. Give a short account of the attempts which have been made to find a common value of the atomic heat for all elements in the solid state.—*Board of Educ.*

56. The equivalent of carbon in methane is 3, in ethylene 6, in acetylene 12. On what considerations is the atomic weight of carbon fixed as 12?—*St. Andrews Univ.*

57. How would you distinguish between (a) a soluble iodide and a soluble bromide; (b) graphite and iodine; (c) nitrous oxide and oxygen; (d) a ferrous and a ferric salt? Give equations.—*St. Andrews Univ.*

58. Describe the methods adopted and the results obtained in the study of the limited oxidation of methane and its homologues.—*Board of Educ.*

59. Sketch the history of the recognition of the property of "radioactivity," and give an account of the isolation of radium salts. Describe exactly with diagrams how you would investigate the nature of the radiation from a given mineral.—*Board of Educ.*

60. Give a short account of the metallic amines. What is known regarding their constitution?—*St. Andrews Univ.*

61. Write a brief essay on valency, taking into account the existence of "molecular" compounds.—*London Univ.*

62. Give an account of the principal ammoniacal bases or amines containing cobalt or platinum, and explain Werner's theory of their constitution.—*Board of Education.*

63. Platinum chloride was added to a solution of ammonium chloride, and the resulting precipitate, after ignition, left 1.7 grams of platinum. What weight of ammonium chloride was present in the solution? ($Pt = 195$, $N = 14$, $H = 1$, $Cl = 35.5$).—*Board of Educ.*

64. Give the formulae of the following substances, state whether each is soluble or insoluble in water, and suggest pairs of solutions which would give the insoluble substances as precipitates: (a) calcium chloride; (b) barium carbonate; (c) aluminium hydroxide; (d) phosphorus chloride; (e) phosphoric chloride; (f) copper sulphide; (g) ammonium sulphate; (h) strontium nitrate.—*Univ. Toronto.*

65. Describe two experimental proofs of each of the following statements: (a) the oxygen, nitrogen, and carbon dioxide in the atmosphere are not chemically combined; (b) the oxygen, nitrogen, and hydrogen in nitric acid are chemically combined.—*Cambridge Senior Locals.*

66. State exactly how you would separately prepare from phosphorus specimens of phosphorous acid, orthophosphoric acid, and hypophosphorous acid. How may these acids be recognized qualitatively? What structural formula would you ascribe to orthophosphoric acid, and for what reasons?—*Institute of Chem.*

67. If the structural formula of phosphorous acid be correctly represented by $P = (OH)_3$, how would you expect this acid to decompose on heating?

68. What reactions take place when (a) silver nitrate solution is acted upon by (i.) arseniuretted hydrogen, (ii.) antimonuretted hydrogen, and (iii.) chlorine; (b) when sulphuretted hydrogen acts on ferric chloride solution; and (c) when solutions of the following are brought in contact: (1) silver nitrate and sodium phosphate, (2) potassium iodide and copper sulphate, (3) ferrous sulphate and nitric acid?—*Cape Univ.*

69. Give examples of monatomic, diatomic, triatomic, and tetratomic molecules of elementary substances, and state how their construction has been arrived at.—*Science and Art Dept.*

70. What are the natural sources of arsenic? Describe how the element is obtained from its ores, and mention some of the uses to which arsenic and its compounds are put. How could you test a specimen of iron pyrites for arsenic?—*Cambridge Univ.*

71. Discuss the action of zinc on aqueous solutions of the following: sodium hydroxide, nitrous acid, sulphurous acid, cupric chloride, arsenic acid.—*Institute of Chem.*

72. Describe and explain the changes which take place in the following experiments: (a) carbon dioxide is passed into lime-water; (b) chlorine is passed into water containing mercuric oxide in suspension; (c) sulphurous acid gas is passed into a solution of iodine; (d) sulphurous acid gas is passed into a solution of hydrogen iodate; (e) potassium permanganate solution is added to a solution of

hydrogen peroxide in dilute sulphuric acid ; (f) phosphorus is boiled with baryta water.—*Owens Coll.*

73. On what experimental evidence is it believed that acids, bases, and salts are more or less ionized in aqueous solution ? Explain from this point of view (a) the alkalinity of a sodium carbonate solution ; (b) the acidity of a ferrous chloride solution ; (c) the action of ammonium chloride in preventing the precipitation of magnesium salts by ammonia.—*Madras Univ.*

74. Coal gas of the composition given below is burnt in air :—

H_2	CH_4	C_2H_2	C_2H_6	C_6H_6	CO	CO_2	CS_2	N_2
48.0,	26.0,	2.0,	3.0,	1.0,	15.0,	2.0,	0.1,	2.7 per cent.

Name the products, and give the quantities of each produced per litre at N.T.P. of gas burnt, stating your results in c.c. measured at 100° and 760 mm.—*Institute of Chem.*

75. Describe an electrolytic method for preparing sodium. Show how the modern applications of electrolysis have (a) aided in the production of cyanides, (b) reduced the profits of the Leblanc soda process, (c) aided in the production of bleaching liquids.—*Cape Univ.*

76. In 1813 Davy wrote : " I have given an account of an experiment on the combustion of potassium in silicated fluoric acid gas " (SiF_4), " in which the gas was absorbed, and a fawn-coloured substance formed, which effervesced with water, and left, after its action on that fluid, a residuum which burnt when heated with oxygen, reproducing silicated fluoric acid gas ; and I concluded from the phenomena that the acid gas was decomposed in the process, that oxygen was probably separated from it by potassium, and that the combustible substance was a compound of the siliceous and fluoric bases." Elucidate, as far as you can, what happened in the experiment described, and comment upon Davy's explanation of it, and on the terminology he uses.—*Owens Coll.*

77. Pieces of bright metallic lead are placed (a) in distilled water ; (b) dilute hydrochloric acid ; (c) dilute sulphuric acid ; (d) water saturated with carbon dioxide ; (e) a chalk water of about 15° hardness ; give an account of the changes which you would expect to observe or be able to trace.—*Institute of Chem.*

78. Each of the following substances is heated in a stream of hydrogen : PbO , SnO_2 , Fe_2O_3 , TiO_2 , Cr_2O_3 , Mn_2O_3 , Al_2O_3 , P_2O_5 , B_2O_3 , ZnO , BaO . State what occurs in each case, giving equations.—*Institute of Chem.*

79. What is the effect of heat on the following substances : arsenic acid, phosphorus pentachloride, lead nitrate, auric chloride, lead dioxide ?—*London Univ.*

80. What happens when the following substances are warmed with concentrated nitric acid : tin, antimony, aluminium, arsenious oxide, phosphorus, hydrogen chloride, ammonia ?—*London Univ.*

81. A graduated gas cylinder is supported vertically in a mercury trough so that the upper end of the tube is distant a mm. from the level of the mercury, and the level of the mercury inside the cylinder is distant b mm. from the outside level. The tube must be depressed c mm. in order to make the level of the mercury inside and outside the same. What is the pressure of the atmosphere outside ? If the height of the tube from the level of the mercury is 300 mm., the difference in the levels of the mercury in the two vessels, 152 mm., and the levels of mercury inside and outside are the same when the tube is depressed 187.6 mm., show that the height of the barometer is 760 mm. Answer : $b(a-b)/(c-3)$.

82. Show that a deviation from the law of persistence of weight would be inconsistent with the law of persistence of energy. Hint : If an atom of weight w , when combining with another atom of weight w' does not produce a molecule of weight $w + w'$, and the weight of the molecule be greater than that of the two constituent atoms, it would suffice to let the molecule fall some convenient distance, and raise this again while the atoms are uncombined to gain a little energy.

83. On heating strongly 0.406 gram of potassium chlorate, 139 c.c. of oxygen were obtained, and the potassium chloride remaining weighed 0.247 gram. The oxygen was measured over water at a temperature of 27° , at which temperature the vapour pressure of water is 2.66 cm. The barometric height was 71.14 cm. Find the value for the molecular weight of oxygen which may be deduced from these data.

84. Explain the following quotation : " When the term ' salt ' is used in everyday life, it refers naturally to sodium chloride, but to a chemist a salt may not

only be a chloride, but a sulphide, a sulphate, or one of a number of similar compounds."—W. Segerblom, *First Year Chemistry*, Exeter, 1909.

85. T. Thomson (*A System of Chemistry*, London, 1817) said: "The atomic theory seems to me to present an insuperable objection to the opinion advocated by Berthollet, that mass produces an effect upon chemical combinations and decompositions." Show that there is no difficulty in reconciling Berthollet's generalization (p. 119) with Dalton's atomic theory.

86. To what constituents do acids owe their characteristic properties? Does this substance impart these properties to all its compounds? What is the theory advanced to explain the facts embraced in your answer? How do we explain the fact that some acids are stronger than others? Name some other facts which the above-mentioned theory explains.—*American Coll.*

87. State, with explanations, how electrolytes differ from non-electrolytes as regards the depression of freezing point or the osmotic pressure exerted by equimolecular solutions, and explain how the degree of ionization of a salt in solution might be calculated from a determination of osmotic pressure. a grams of a non-electrolyte are made up to 100 c.c. with water at 15°C . The osmotic pressure is found to be 500 mm. Express in factors the molecular weight of the substance.—*Cape Univ.*

88. What is the difference between the chemical equivalent and the electrochemical equivalent? (a) The chemical equivalents of zinc, aluminium, and (ferrous) iron are respectively 32.5, 9, and 28. What are the respective electrochemical equivalents of the elements? Distinguish between "unit quantity of electricity" and "unit current of electricity." (b) How many units of current will deposit 0.01118 gm. of silver per minute? *Answers.* (a) The electrochemical equivalent of zinc is 32.5, aluminium 9, and ferrous iron 28; the electrochemical equivalents in coulombs are respectively 0.00034, 0.000095, and 0.00029. (b) $\frac{1}{2}$ amp.

89. The durability of galvanized iron is, under ordinary circumstances, much greater than that of tin plates (iron coated with tin). Explain this.—*New Zealand Univ.*

90. If hydrogen sulphide be passed into a mixture of iodine and water, sulphur is deposited, and hydrogen iodide is formed, the iodine may be in turn displaced from combination by bromine, and the bromine by chlorine. What weight of each of these elements would be required to displace the sulphur contained in 100 grms. of hydrogen sulphide? ($\text{S} = 23$, $\text{I} = 127$, $\text{Br} = 80$, $\text{Cl} = 35.5$).—*London Univ.*

91. What do you understand by positive ions, negative ions, and complex ions? Give examples. Give examples of metals (two at least in each case): (a) which form more than one simple positive ion; (b) which form only one oxide or hydroxide, and this oxide or hydroxide, although usually giving a simple positive ion, sometimes forms part of a negative ion; (c) which form more than one oxide, and in certain salts form part of a negative ion; (d) which in some parts form part of a complex positive ion. Give the formula of the ion in each case.—*Sydney Univ.*

92. The analysis of a potassium hydrogen arsenomolybdate by C. Friedheim (1892) gave the following data: 0.4328 gm. analyzed for potash gave 0.1138 gm. of platinum; for arsenic, 0.78819 gm. of substance gave 0.3418 gm. $\text{Mg}_3\text{P}_2\text{O}_7$; 0.5038 gm. weighed 0.0587 gm. after ignition. The MoO_3 was determined by difference. Show that these numbers correspond with 12.67 per cent. of K_2O ; 32.41 per cent. of As_2O_5 ; 11.65 per cent. H_2O ; and 43.27 per cent. MoO_3 ; and hence agree with the formula $\text{K}_2\text{O} \cdot \text{As}_2\text{O}_5 \cdot 2\text{MoO}_3 \cdot 5\text{H}_2\text{O}$.

93. The element manganese is commonly classed as a metal. Point out how this is justified. There are, however, some compounds of manganese, the study of which would lead to its classification with other elements certainly not metals. Give an account of any of these you can, and say what comparisons they suggest to you.

94. When 0.2041 gm. of anhydrous ferrous chloride was exposed to ammonia gas at 18.52°C . H. Wolfram (1913) found that 235.1 c.c. were absorbed. Barometer 746 mm. Show that this very nearly corresponds with the formation of $\text{FeCl}_2 \cdot 6\text{NH}_3$.

95. Show that the heat of formation of phosphorus nitride, P_3N_5 , from red phosphorus is nearly 70.4 cal., given the heats of combustion of the nitride P_3N_5 to nitrogen and phosphorus pentoxide is 474.7 cal., and of red phosphorus 370.8 cal.; and further, that the heat of formation of the nitride from colourless

104. How would you distinguish between: (a) nitrous oxide and oxygen, (b) ethylene and acetylene, (c) carbon tetrachloride and silicon tetrachloride?—*Board of Educ.*

105. What is meant by the terms: (a) "hard water," (b), "soft water," (c) "temporary hardness," (d) "permanent hardness"? What is the cause of (a) "temporary hardness," (b) "permanent hardness," of water, and how may each of these kinds of "hardness" be removed?—*Board of Educ.*

106. What are the principal ores of chromium? How is potassium dichromate usually prepared commercially? Starting from potassium dichromate, how would you prepare (a) chromium sulphate, (b) chromium oxychloride, (c) potassium chromate?—*Board of Educ.*

107. State and explain carefully the law of Mass Action, supplying at least one illustration of its application to gases, one to solutions, and one to a reaction between a gas and a solid.—*Board of Educ.*

108. Give, with equations, the reactions by which you would detect in the presence of each other: a sulphide, sulphite, sulphate and thiosulphate; a chloride, bromide, nitrite, and nitrate.—*Board of Educ.*

109. Explain the terms (a) enantiotropic, (b) transition point, and describe some method by which the latter can be determined.—*Board of Educ.*

110. Outline the conditions under which persulphuric acid may be formed. What are the chief properties of the acid and of its salts? What formula would you assign to the acid, and why?—*Board of Educ.*

111. What is the action on a solution of copper sulphate of the following substances, also in solution: (a) caustic soda, (b) potassium hypophosphite, (c) potassium cyanide, (d) potassium iodide, (e) sulphur dioxide, and (f) potassium thiocyanate? Give equations. Describe a volumetric method of determining copper based on any of these reactions.—*Board of Educ.*

112. Give an account of the production of "cyanide" from either (a) ammonia, or (b) thiocyanate (from gas liquor). For what industrial purposes is "cyanide" required? How may (a) cyanogen, (b) Prussian blue, (c) carbon monoxide be obtained from potassium cyanide?—*Board of Educ.*

113. Describe in sufficient detail the modern "lead chamber" process for the manufacture of sulphuric acid, and discuss the theory of the reactions involved.—*Board of Educ.*

114. (a) Describe the properties of radium compounds. (b) In what respects does radium differ from most other elements?—*Harvard Univ.*

115. Describe the process of extraction of radium from pitchblende.—*Calcutta Univ.*

116. What led Crookes to propound his theory of meta-elements? How far is this theory tenable in the light of modern research?—*Calcutta Univ.*

117. Give a short account of the chemistry of radium.—*Punjab Univ.*

118. How, if at all, have the discoveries of radioactivity influenced chemical theory; or in other words, give a sketch of the theory of radioactivity and its relation to the atomic theory.—*Punjab Univ.*

119. Give an account of recent work with radium and the degradation of other elements.—*Cape Univ.*

120. The price of radium salts in a catalogue (1913) is quoted at the rate of £12,500 per gram for both radium chloride and radium bromide. Other things being equal, which do you consider the more advantageous mode of purchase?

121. Name the principal "impurities" which are usually present in the air of towns. How could you determine the amount of any two of these substances present in 100 volumes of air? How does town air differ from country air, and how may these differences be accounted for?—*Board of Educ.*

122. Under what circumstances, if any, does water react chemically with the following substances: Sodium peroxide, sodium amalgam, calcium carbide, silicon tetrafluoride, and phosphorus pentachloride? Explain briefly the nature of the chemical changes taking place in each case.—*Board of Educ.*

123. Describe fully an experiment by means of which it may be shown that sulphur dioxide gas contains its own volume of oxygen. Explain clearly the reactions, if any, which occur when this gas is passed into each of the following liquids: (a) lime-water; (b) caustic soda solution, (c) an acidified solution of potassium permanganate, (d) potassium iodate solution, (e) fuming nitric acid.—*Board of Educ.*

124. Starting from lead sulphide, describe how you would prepare (a) lead, (b) litharge, (c) "red lead," (d) white lead? How would you detect the presence

of lead peroxide and of lead chloride in a mixture of these two substances ?—*Board of Educ.*

125. How do you explain the following : (a) A solution of sodium silicate is readily decomposed by weak acids, but sodium sulphate heated at a high temperature with silica forms sodium silicate ? (b) Barium peroxide heated at a certain temperature under reduced pressure forms barium monoxide with the liberation of oxygen, whereas barium monoxide heated at the same temperature under increased pressure combines with oxygen forming barium dioxide ?—*Board of Educ.*

126. Make a comparison of the compounds of carbon and silicon, so as to reveal the relationship existing between them. What suggestion can be made to account for the fact that carbon dioxide is a gas, whereas silica is a solid of high melting point ? What abnormalities are found in connection with the specific heats of the two elements ?—*Board of Educ.*

127. Explain how the valency of an element is connected with the position of the element in the Periodic table. What are the reasons for the positions assigned to (a) hydrogen, (b) manganese or copper, in the table ?—*Board of Education.*

128. What reactions take place when atmospheric nitrogen is passed over (a) calcium, (b) calcium carbide, (c) barium carbide, (d) calcium hydride ? What action, if any, has water on (1) the substances named, (2) the products obtained ?—*Board of Educ.*

129. What do you understand by the term "relative migration velocity" ? In what way is the relative migration velocity connected with the molecular conductivity of an electrolyte ? Describe an experiment by which such a velocity can be determined.—*Board of Educ.*

130. What is meant by the term "hydrolysis" ? Explain clearly the various factors which influence the degree of hydrolytic dissociation. State and explain what occurs when potassium cyanide, ferric chloride, and borax are separately dissolved in water ; and what happens when ammonium carbonate is added to a solution of aluminium sulphate.—*Board of Educ.*

131. Explain what occurs when carbon monoxide reacts with (a) ammoniacal cuprous chloride, (b) chlorine, (c) nickel, (d) iron. Describe the properties of the several products and the conditions under which they are obtained. Have any of these products received technical application, and for what purposes, if any, is each applied ?—*Board of Educ.*

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